# HL Paper 2

The periodic table shows the relationship between electron configuration and the properties of elements and is a valuable tool for making predictions in chemistry.

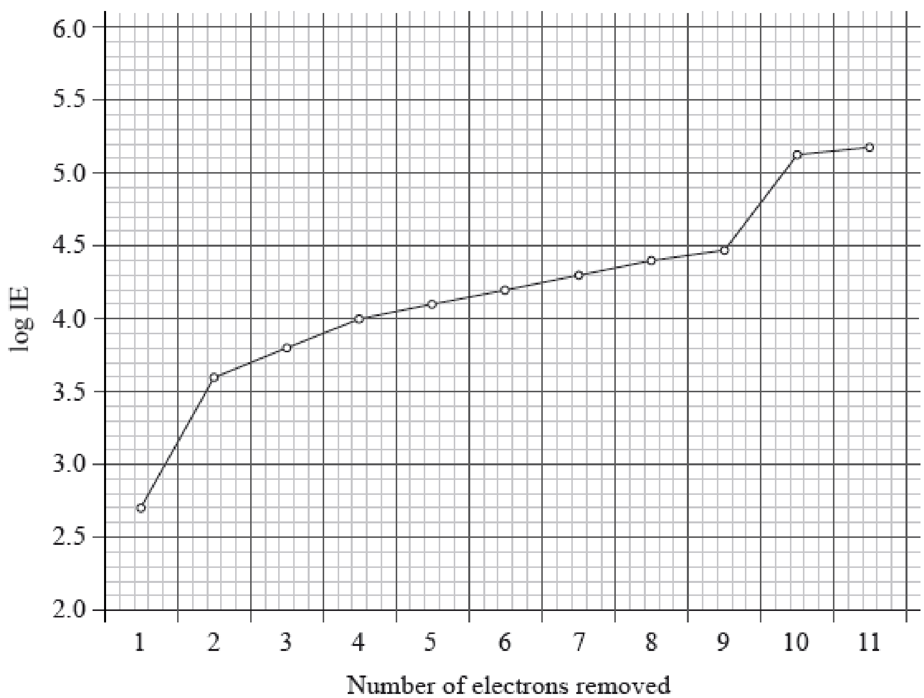
The ten elements in the first-row d-block have characteristic properties and many uses.

b.i. Define the term *electronegativity*. [2]

c. (i) Outline **two** reasons why a sodium ion has a smaller radius than a sodium atom. [4]

(ii) Explain why the ionic radius of $P^{3-}$ is **greater** than the ionic radius of $Si^{4+}$.

d. The graph below represents the successive ionization energies of sodium. The vertical axis plots log (ionization energy) instead of ionization energy to allow the data to be represented without using an unreasonably long vertical axis. [4]




State the full electron configuration of sodium and explain how the successive ionization energy data for sodium are related to its electron configuration.

e. (i) Explain why the first ionization energy of aluminium is **lower** than the first ionization energy of magnesium. [4]

(ii) Explain why the first ionization energy of sulfur is **lower** than the first ionization energy of phosphorus.

f.i. State and explain the type of reaction that takes place between $Fe^{3+}$ and $H_2O$ to form $[Fe(H_2O)_6]^{3+}$ in terms of acid-base theories. [2]

f.ii. Explain why $[Fe(H_2O)_6]^{3+}$ is coloured. [3]

f.iii. Outline the economic significance of the use of a catalyst in the Haber process which is an exothermic reaction. [2]

The emission spectrum of an element can be used to identify it.

a.iii Hydrogen spectral data give the frequency of $3.28 \times 10^{15}$ $s^{-1}$ for its convergence limit. [1]

Calculate the ionization energy, in J, for a single atom of hydrogen using sections 1 and 2 of the data booklet.

a.iv Calculate the wavelength, in m, for the electron transition corresponding to the frequency in (a)(iii) using section 1 of the data booklet. [1]

c.iv Deduce any change in the colour of the electrolyte during electrolysis. [1]

c.v. Deduce the gas formed at the anode (positive electrode) when graphite is used in place of copper. [1]

d. Explain why transition metals exhibit variable oxidation states in contrast to alkali metals. [2]

Transition metals:

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

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Alkali metals:

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Copper is a metal that has been used by humans for thousands of years.

a. State the full electron configuration of $^{65}\mathrm{Cu}$. [1]

b. State one difference in the physical properties of the isotopes $^{63}\mathrm{Cu}$ and $^{65}\mathrm{Cu}$ and explain why their chemical properties are the same. [2]

Physical:

Chemical:

c. Describe the bonding in solid copper. [2]

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Bonds can be formed in many ways.

Bonds can be formed in many ways.

The equilibrium for a mixture of $NO_2$ and $N_2O_4$ gases is represented as:

$$2NO_2(g) \rightleftharpoons N_2O_4(g)$$

At 100°C, the equilibrium constant, $K_c$, is 0.21.

b.i. Discuss the bonding in the resonance structures of ozone. [3]

b.ii. Deduce one resonance structure of ozone and the corresponding formal charges on each oxygen atom. [2]

c. The first six ionization energies, in kJ mol$^{-1}$, of an element are given below. [2]

| $IE_1$ | $IE_2$ | $IE_3$ | $IE_4$ | $IE_5$ | $IE_6$ |
|---|---|---|---|---|---|
| 578 | 1816 | 2744 | 11 576 | 14 829 | 18 375 |

Explain the large increase in ionization energy from $IE_3$ to $IE_4$.

d.i. At a given time, the concentration of $NO_2(g)$ and $N_2O_4(g)$ were 0.52 and $0.10 \text{ mol dm}^{-3}$ respectively. [2]

Deduce, showing your reasoning, if the forward or the reverse reaction is favoured at this time.

d.ii. Comment on the value of $\Delta G$ when the reaction quotient equals the equilibrium constant, $Q = K$. [2]

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Magnesium, a reactive metal found in many common minerals, is also an essential nutrient for both plants and animals.

Successive ionization energies of magnesium are given in the table below.

| | First | Second | Third |
|---|---|---|---|
| **Energy required / kJ mol$^{-1}$** | 738 | 1450 | 7730 |

Magnesium metal is mainly used as a component in lightweight alloys, particularly in combination with aluminium and titanium.

Magnesium is usually produced by the electrolysis of molten magnesium chloride.

a. Define the term *first ionization energy*. [2]

b. (i) Explain why the second ionization energy is greater than the first ionization energy. [4]

(ii) Explain why the third ionization energy is much greater than the second ionization energy.

c. Although magnesium is usually found as $Mg^{2+}$ in its compounds, it is possible to use the Born-Haber cycle to investigate the possibility of $Mg^{+}$ being able to form stable compounds. [3]

Use the ionization energy data from part (b), along with the other data provided below, to determine the enthalpy change of formation of MgCl(s). Assume that, because $Mg^{+}$ would be similar in size to $Na^{+}$, MgCl would have a similar lattice enthalpy to NaCl.

Enthalpy of atomization of Mg $+146\ \mathrm{kJ\,mol^{-1}}$

Bond enthalpy in $Cl_2$ $+243\ \mathrm{kJ\,mol^{-1}}$

Electron affinity of Cl $+349\ \mathrm{kJ\,mol^{-1}}$

Lattice enthalpy of NaCl $+790\ \mathrm{kJ\,mol^{-1}}$

d. Consider the lattice enthalpies of $MgF_2$, $MgCl_2$ and $CaCl_2$. List these from the most endothermic to the least endothermic and explain your order. [3]

$\text{Most endothermic} \rightarrow \text{Least endothermic}$

e. Magnesium hydroxide, $Mg(OH)_2$, is only sparingly soluble in water and the equilibrium below exists when excess solid is in contact with a saturated solution. [2]

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^{-}(aq)$$

Outline how the solubility of magnesium hydroxide will vary with pH.

f. (i) Describe the bonding present in magnesium metal. [4]

(ii) Suggest why magnesium is harder than sodium.

(iii) Outline why alloys are generally less malleable than their component metals.

g. (i) Draw a labelled diagram of a suitable apparatus for the electrolysis. [7]

(ii) State equations for the reactions that take place at the electrodes.

Negative electrode (cathode) reaction:

Positive electrode (anode) reaction:

(iii) When dilute aqueous magnesium chloride is used as the electrolyte, the reactions at both electrodes are different. State equations for the reactions that occur in aqueous solution.

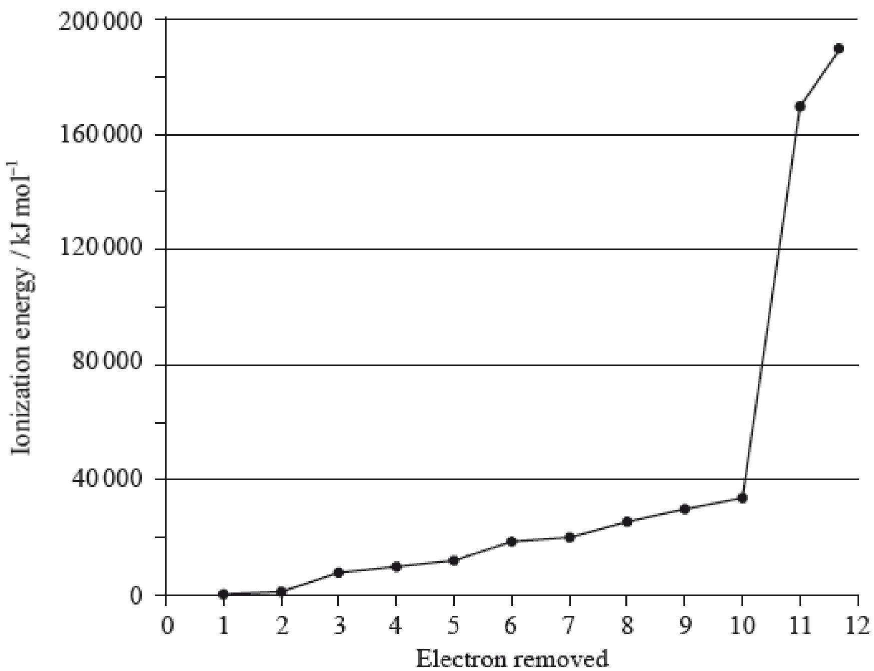
Negative electrode (cathode) reaction:

Positive electrode (anode) reaction:

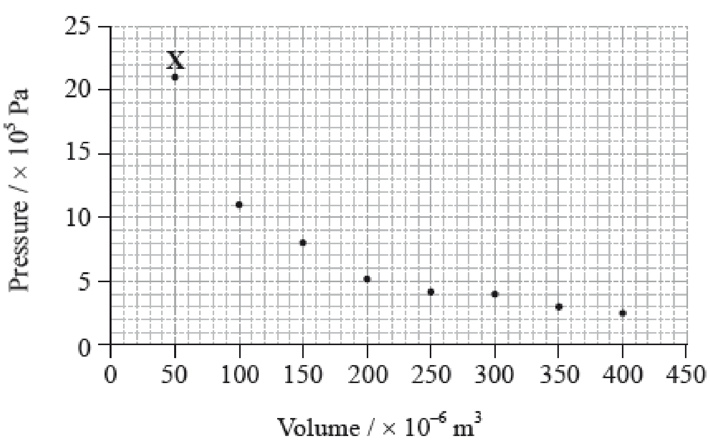
(iv) Outline why magnesium metal is not produced in the electrolysis of aqueous magnesium chloride.

A sample of magnesium contains three isotopes: magnesium-24, magnesium-25 and magnesium-26, with abundances of 77.44%, 10.00% and 12.56% respectively.

A graph of the successive ionization energies of magnesium is shown below.




The graph below shows pressure and volume data collected for a sample of carbon dioxide gas at 330 K.




a. (i) Calculate the relative atomic mass of this sample of magnesium correct to **two** decimal places. [4]

(ii) Predict the relative atomic radii of the three magnesium isotopes, giving your reasons.

b. (i) Explain the increase in ionization energy values from the 3rd to the 8th electrons. [3]

(ii) Explain the sharp increase in ionization energy values between the 10th and 11th electrons.

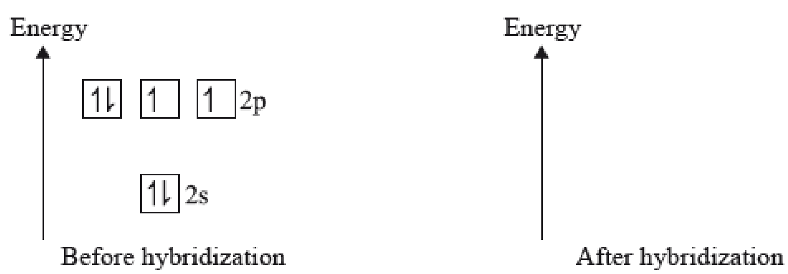
c. (i) Magnesium reacts with oxygen to form an ionic compound, magnesium oxide. Describe how the ions are formed, and the structure and bonding in magnesium oxide. [4]

(ii) Carbon reacts with oxygen to form a covalent compound, carbon dioxide. Describe what is meant by a covalent bond.

(iii) State why magnesium and oxygen form an ionic compound while carbon and oxygen form a covalent compound.

d. (i) Predict the type of hybridization of the carbon and oxygen atoms in $CO_2$. [7]

(ii) Sketch the orbitals of an oxygen atom in $CO_2$ on the energy level diagram provided, including the electrons that occupy each orbital.




(iii) Define the term electronegativity.

(iv) Explain why oxygen has a larger electronegativity than carbon.

e. (i) Draw a best-fit curve for the data on the graph. [4]

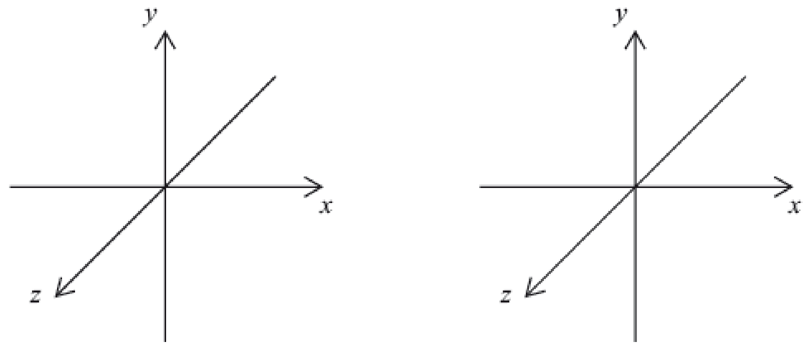
(ii) Use the data point labelled **X** to determine the amount, in mol, of carbon dioxide gas in the sample.

f. (i) Most indicators are weak acids. Describe qualitatively how indicators work. [3]

(ii) Identify a suitable indicator for a titration between a weak acid and a strong base, using Table 16 of the Data Booklet.

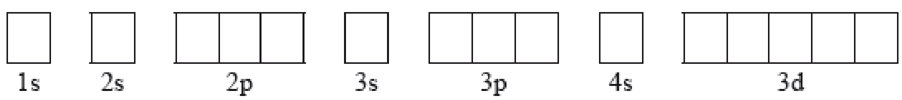
Lithium and boron are elements in period 2 of the periodic table. Lithium occurs in group 1 (the alkali metals) and boron occurs in group 3. Isotopes exist for both elements.

a.iii.The electron configuration of boron is $1s^22s^22p^1$. Draw the shape of an s orbital and a $p_x$ orbital on the axes below. [1]




s orbital $\qquad$ $p_x$ orbital

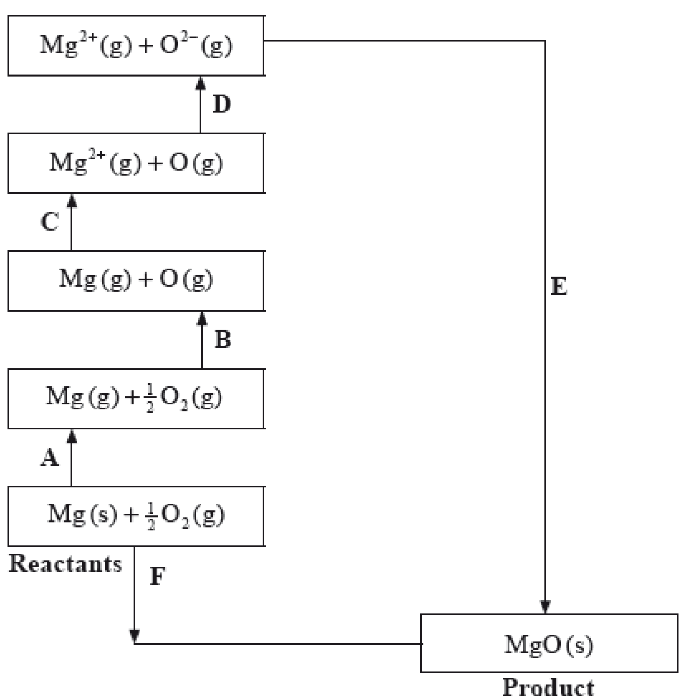
b. (ii) Cobalt is a transition metal. One common ion of cobalt is $Co^{3+}$. Draw the orbital diagram (using the arrow-in-box notation) for the $Co^{3+}$ ion. [5]




(iii) State the other most common ion of cobalt.

(iv) Explain why the complex $[Co(NH_3)_6]Cl_3$ is coloured.

The Born-Haber cycle for MgO under standard conditions is shown below.




The values are shown in the table below.

| Process | Enthalpy change / kJ $mol^{-1}$ |
|---|---|
| **A** | +150 |
| **B** | +248 |
| **C** | +736 + (+1450) |
| **D** | −142 + (+844) |
| **E** | |
| **F** | −602 |

a.i. Identify the processes represented by **A**, **B** and **D** in the cycle. [3]

a.ii. Define the enthalpy change, **F**. [2]

a.iii. Determine the value of the enthalpy change, **E**. [2]

a.iv. Define the enthalpy change **C** for the first value. Explain why the second value is significantly larger than the first. [4]

a.v. The inter-ionic distance between the ions in NaF is very similar to that between the ions in MgO. Suggest with a reason, which compound has the higher lattice enthalpy value. [2]

b.i. The standard enthalpy change of three combustion reactions is given below in kJ. [4]

$$\begin{array}{ll} 2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l) & \Delta H^\ominus = -3120 \\ 2H_2(g) + O_2(g) \rightarrow 2H_2O(l) & \Delta H^\ominus = -572 \\ C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l) & \Delta H^\ominus = -1411 \end{array}$$

Based on the above information, calculate the standard change in enthalpy, $\Delta H^\ominus$, for the following reaction.

$$C_2H_6(g) \rightarrow C_2H_4(g) + H_2(g)$$

b.ii. Predict, stating a reason, whether the sign of $\Delta S^\ominus$ for the above reaction would be positive or negative. [2]

b.iii. Discuss why the above reaction is non-spontaneous at low temperature but becomes spontaneous at high temperatures. [2]

b.iv. Using bond enthalpy values, calculate $\Delta H^\ominus$ for the following reaction. [3]

$$C_2H_6(g) \rightarrow C_2H_4(g) + H_2(g)$$

b.v. Suggest with a reason, why the values obtained in parts (b) (i) and (b) (iv) are different. [1]

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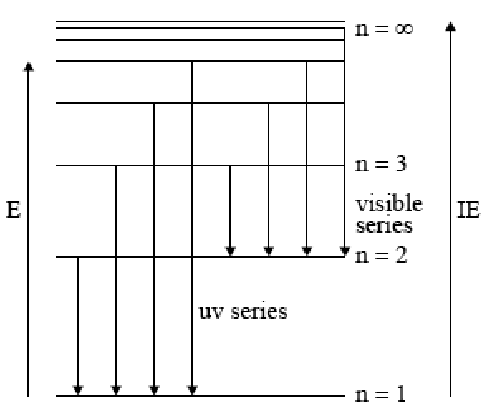

On the above diagram, draw the line that corresponds to the first ionization energy of hydrogen and explain your reasoning.

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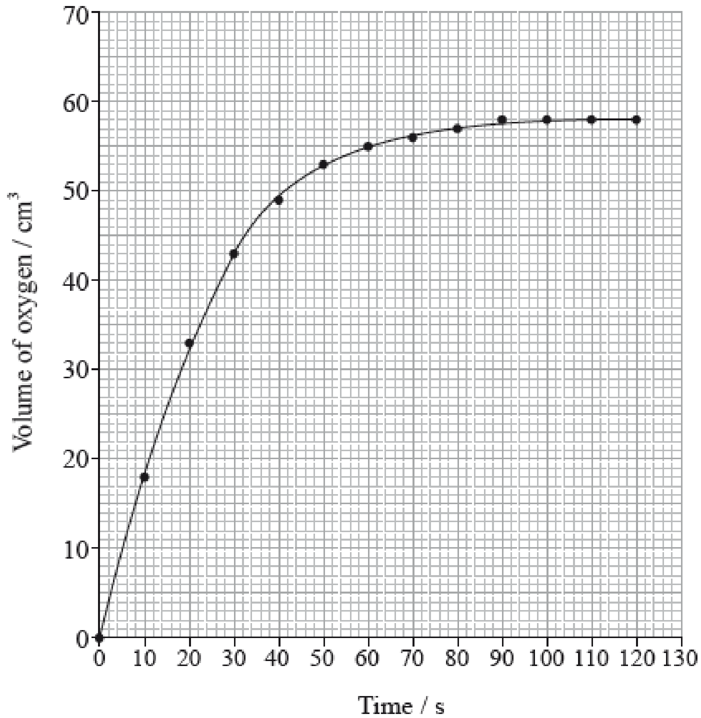
Hydrogen peroxide decomposes according to the equation below.

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

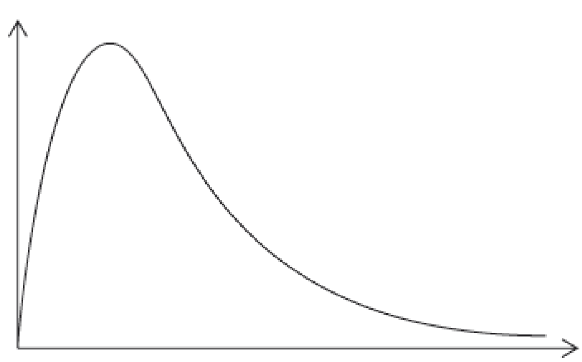
The rate of the decomposition can be monitored by measuring the volume of oxygen gas released. The graph shows the results obtained when a solution of hydrogen peroxide decomposed in the presence of a CuO catalyst.




a.i. Outline how the initial rate of reaction can be found from the graph. [2]

a.ii. Explain how and why the rate of reaction changes with time. [3]

b. A Maxwell-Boltzmann energy distribution curve is drawn below. Label both axes and explain, by annotating the graph, how catalysts increase the rate of reaction. [3]



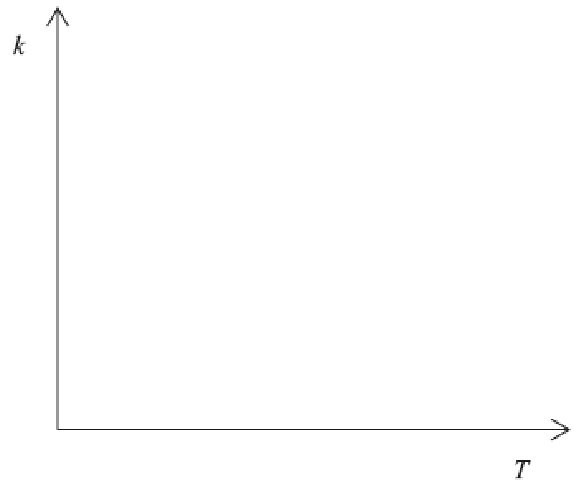
c. (i) In some reactions, increasing the concentration of a reactant does not increase the rate of reaction. Describe how this may occur. [2]

(ii) Consider the reaction

$$2A + B \rightarrow C + D$$

The reaction is first order with respect to **A**, and zero order with respect to **B**. Deduce the rate expression for this reaction.

d. Sketch a graph of rate constant $(k)$ versus temperature. [1]




e. Hydrochloric acid neutralizes sodium hydroxide, forming sodium chloride and water. [9]

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) \quad \Delta H^\ominus = -57.9 \text{ kJ mol}^{-1}$$

(i) Define *standard enthalpy change of reaction*, $\Delta H^\ominus$.

(ii) Determine the amount of energy released, in kJ, when $50.0 \text{ cm}^3$ of $1.00 \text{ mol dm}^{-3}$ sodium hydroxide solution reacts with $50.0 \text{ cm}^3$ of $1.00 \text{ mol dm}^{-3}$ hydrochloric acid solution.

(iii) In an experiment, 2.50 g of solid sodium hydroxide was dissolved in $50.0 \text{ cm}^3$ of water. The temperature rose by 13.3 °C. Calculate the standard enthalpy change, in $\text{kJ mol}^{-1}$, for dissolving one mole of solid sodium hydroxide in water.

$$NaOH(s) \rightarrow NaOH(aq)$$

(iv) Using relevant data from previous question parts, determine $\Delta H^\ominus$, in $\text{kJ mol}^{-1}$, for the reaction of solid sodium hydroxide with hydrochloric acid.

$$NaOH(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$$

f. (i) Zinc is found in the d-block of the periodic table. Explain why it is not considered a transition metal. [5]

(ii) Explain why $Fe^{3+}$ is a more stable ion than $Fe^{2+}$ by reference to their electron configurations.

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Calcium carbide, $CaC_2$, is an ionic solid.

a. Describe the nature of ionic bonding. [1]

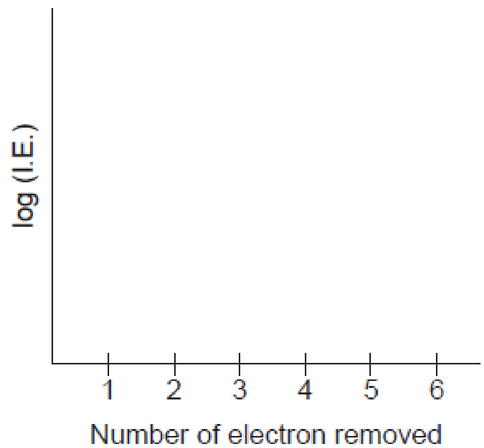
b. Describe how the relative atomic mass of a sample of calcium could be determined from its mass spectrum. [2]

c. When calcium compounds are introduced into a gas flame a red colour is seen; sodium compounds give a yellow flame. Outline the source of the colours and why they are different. [2]

d.i. Suggest **two** reasons why solid calcium has a greater density than solid potassium. [2]

d.ii. Outline why solid calcium is a good conductor of electricity. [1]

e. Sketch a graph of the first six ionization energies of calcium. [2]




f. Calcium carbide reacts with water to form ethyne and calcium hydroxide. [1]

$$CaC_2(s) + H_2O(l) \rightarrow C_2H_2(g) + Ca(OH)_2(aq)$$

Estimate the pH of the resultant solution.

g.i. Describe how sigma ($\sigma$) and pi ($\pi$) bonds are formed. [2]

sigma ($\sigma$):

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

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pi ($\pi$):

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

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g.ii. Deduce the number of $\sigma$ and $\pi$ bonds in a molecule of ethyne. [1]

sigma ($\sigma$):

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

pi ($\pi$):

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

Iron has three main naturally occurring isotopes which can be investigated using a mass spectrometer.

d. State the full electronic configurations of a Cu atom and a $Cu^{+}$ ion. [2]

Cu:

$Cu^{+}$:

e. Explain the origin of colour in transition metal complexes and use your explanation to suggest why copper(II) sulfate, $CuSO_4$(aq), is blue, but zinc sulfate, $ZnSO_4$(aq), is colourless. [4]

f. $Cu^{2+}(aq)$ reacts with ammonia to form the complex ion $[Cu(NH_3)_4]^{2+}$. Explain this reaction in terms of an acid-base theory, and outline how the bond is formed between $Cu^{2+}$ and $NH_3$. [3]

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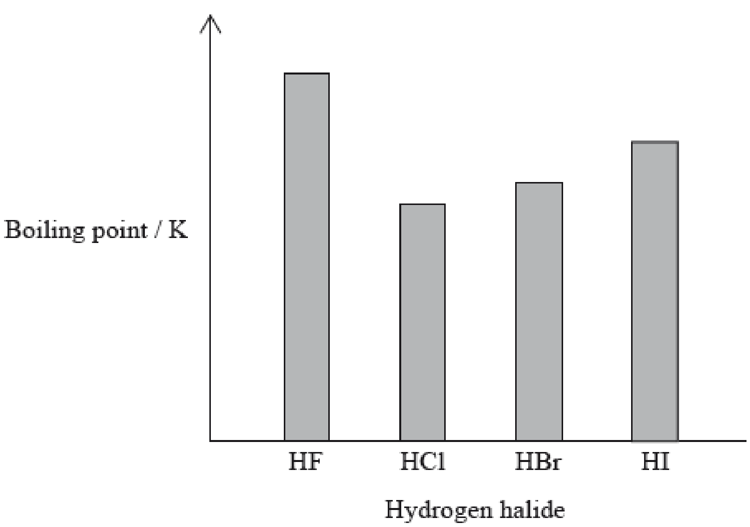
The oxides and chlorides of period 3 elements exhibit periodicity.

Chlorine gas, $Cl_2(g)$, is bubbled through separate solutions of aqueous bromine, $Br_2(aq)$, and potassium bromide, $KBr(aq)$.

The hydrogen halides do not show perfect periodicity. A bar chart of boiling points shows that the boiling point of hydrogen fluoride, HF, is much higher than periodic trends would indicate.




Transition metals form complex ions which are usually coloured.

a. (i) State the changes in the acid-base nature of the oxides across period 3 (from $Na_2O$ to $Cl_2O_7$), including equations for the reactions of $Na_2O$ and $SO_3$ with water. [7]

(ii) State whether or not molten aluminium chloride, $Al_2Cl_6$, and molten aluminium oxide, $Al_2O_3$, conduct electricity. Explain this behaviour in terms of the structure and bonding of the two compounds.

(iii) State the equation for the reaction of $Cl_2$ with water.

b. (i) Predict any changes that may be observed in each case. [4]

$Br_2(aq)$:

$KBr(aq)$:

(ii) State the half-equations for the reactions that occur.

c. (i) Explain why the boiling point of HF is much higher than the boiling points of the other hydrogen halides. [3]

(ii) Explain the trend in the boiling points of HCl, HBr and HI.

d.i. State the full electron configurations of Cr and $Cr^{3+}$. [2]

Cr:

$Cr^{3+}$:

d.ii. $Cr^{3+}$ ions and water molecules bond together to form the complex ion $[Cr(H_2O)_6]^{3+}$. [3]

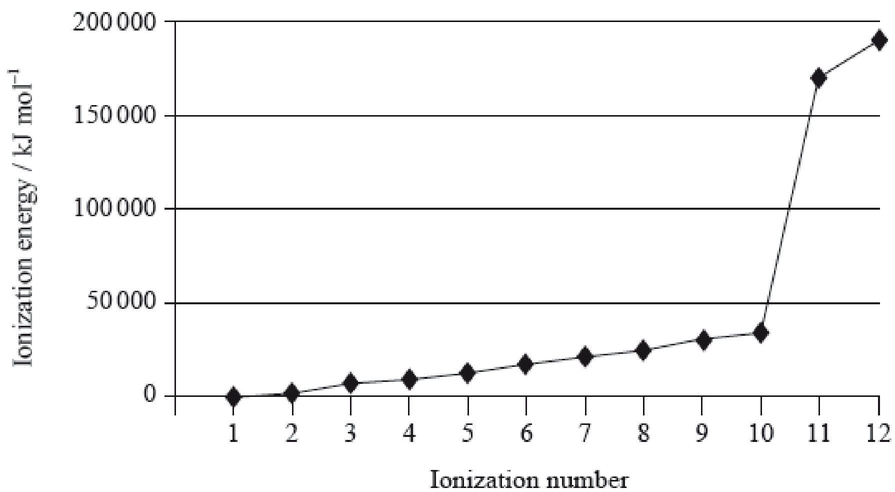
Describe how the water acts and how it forms the bond, identifying the acid-base character of the reaction.

d.iii. Explain why the $[Cr(H_2O)_6]^{3+}$ ion is coloured. [3]

d.iv. Outline, including a relevant equation, whether the $[Cr(H_2O)_6]^{3+}$ ion is acidic, basic or neutral. [1]

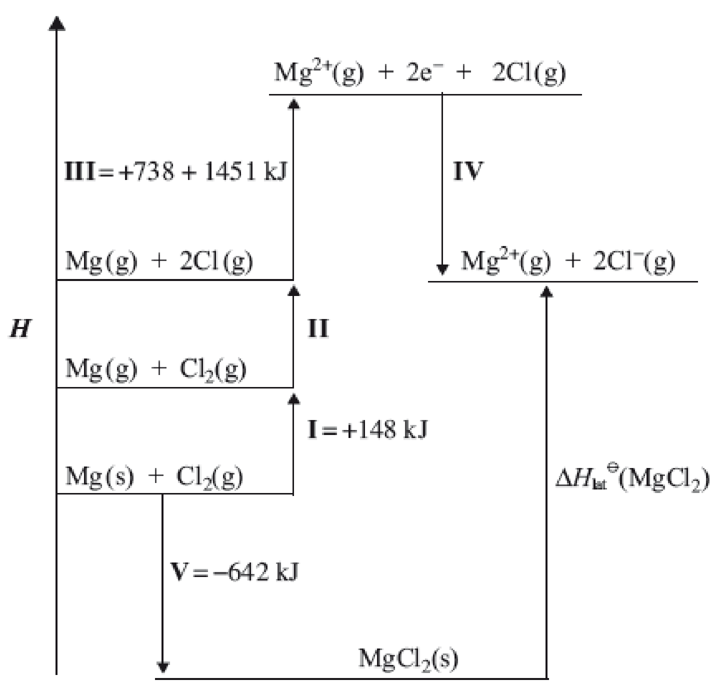
e. Explain how the number of electrons in the outer main energy level of phosphorus, P, can be determined using the data of successive ionization energies. [2]

Magnesium is the eighth most abundant element in the earth's crust. The successive ionization energies of the element are shown below.




Magnesium can be produced from the electrolysis of molten magnesium chloride, $MgCl_2$.

The lattice enthalpy of magnesium chloride can be calculated from the Born-Haber cycle shown below.




a. (i) Define the term *first ionization energy* and state the equation for the first ionization of magnesium. [8]

(ii) Explain the general increase in successive ionization energies of the element.

(iii) Explain the large increase between the tenth and eleventh ionization energies.

b. (i) Explain how molten magnesium chloride conducts an electric current. [5]

(ii) Identify the electrode where oxidation occurs during electrolysis of molten magnesium chloride and state an equation for the half-reaction.

(iii) Explain why magnesium is not formed during the electrolysis of aqueous magnesium chloride solution.

c. (i) Identify the enthalpy changes labelled by **I** and **V** in the cycle. [10]

(ii) Use the ionization energies given in the cycle above and further data from the Data Booklet to calculate a value for the lattice enthalpy of magnesium chloride.

(iii) The theoretically calculated value for the lattice enthalpy of magnesium chloride is +2326 kJ. Explain the difference between the theoretically calculated value and the experimental value.

(iv) The experimental lattice enthalpy of magnesium oxide is given in Table 13 of the Data Booklet. Explain why magnesium oxide has a higher lattice enthalpy than magnesium chloride.

d. (i) State whether aqueous solutions of magnesium oxide and magnesium chloride are acidic, alkaline or neutral. [2]

(ii) State an equation for the reaction between magnesium oxide and water.

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Magnesium is a group 2 metal which exists as a number of isotopes and forms many compounds.

c. Magnesium ions produce no emission or absorption lines in the visible region of the electromagnetic spectrum. Suggest why most magnesium compounds tested in a school laboratory show traces of yellow in the flame. [1]

d. (i) Explain the convergence of lines in a hydrogen emission spectrum. [2]

(ii) State what can be determined from the frequency of the convergence limit.

i. Magnesium chloride can be electrolysed. [5]

(i) Deduce the half-equations for the reactions at each electrode when **molten** magnesium chloride is electrolysed, showing the state symbols of the products. The melting points of magnesium and magnesium chloride are 922K and 987K respectively.

| |
|---|
| Anode (positive electrode):<br><br>. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .<br><br>Cathode (negative electrode):<br><br>. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . |

(ii) Identify the type of reaction occurring at the cathode (negative electrode).

(iii) State the products when a very **dilute** aqueous solution of magnesium chloride is electrolysed.

| |
|---|
| Anode (positive electrode):<br><br>. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .<br><br>Cathode (negative electrode):<br><br>. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . |

j. Standard electrode potentials are measured relative to the standard hydrogen electrode. Describe a standard hydrogen electrode. [2]

k. A magnesium half-cell, $Mg(s)/Mg^{2+}(aq)$, can be connected to a copper half-cell, $Cu(s)/Cu^{2+}(aq)$. [4]

(i) Formulate an equation for the spontaneous reaction that occurs when the circuit is completed.

(ii) Determine the standard cell potential, in V, for the cell. Refer to section 24 of the data booklet.

(iii) Predict, giving a reason, the change in cell potential when the concentration of copper ions increases.

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Titanium and vanadium are consecutive elements in the first transition metal series.

$\mathrm{TiCl_4}$ reacts with water and the resulting titanium(IV) oxide can be used as a smoke screen.

a. Describe the bonding in metals. [2]

b. Titanium exists as several isotopes. The mass spectrum of a sample of titanium gave the following data: [2]

| Mass number | % abundance |
|---|---|
| 46 | 7.98 |
| 47 | 7.32 |
| 48 | 73.99 |
| 49 | 5.46 |
| 50 | 5.25 |

Calculate the relative atomic mass of titanium to two decimal places.

c. State the number of protons, neutrons and electrons in the $^{48}_{22}\mathrm{Ti}$ atom. [1]

Protons: ..........

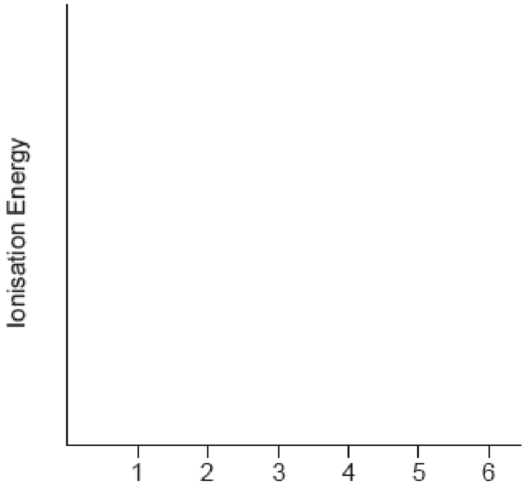
Neutrons: ..........

Electrons: ..........

d.i. State the full electron configuration of the $^{48}_{22}\mathrm{Ti}^{2+}$ ion. [1]

d.ii. Suggest why the melting point of vanadium is higher than that of titanium. [1]

d.iii. Sketch a graph of the first six successive ionization energies of vanadium on the axes provided. [1]




d.iv. Explain why an aluminium-titanium alloy is harder than pure aluminium. [2]

e. Describe, in terms of the electrons involved, how the bond between a ligand and a central metal ion is formed. [1]

f. Outline why transition metals form coloured compounds. [4]

g.i. State the type of bonding in potassium chloride which melts at 1043 K. [1]

g.ii. A chloride of titanium, $\mathrm{TiCl_4}$, melts at 248 K. Suggest why the melting point is so much lower than that of KCl. [1]

h.i. Formulate an equation for this reaction. [2]

h.ii. Suggest one disadvantage of using this smoke in an enclosed space. [1]

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Tin(II) chloride is a white solid that is commonly used as a reducing agent.

a. (i) State why you would expect tin(II) chloride to have a similar lattice enthalpy to strontium chloride, using section 9 of the data booklet. [4]

(ii) Calculate the molar enthalpy change when strontium chloride is dissolved in water, using sections 18 and 20 of the data booklet.

(iii) Tin(II) chloride reacts with water to precipitate the insoluble basic chloride, Sn(OH)Cl.

$$SnCl_2(aq) + H_2O(l) \rightleftharpoons Sn(OH)Cl(s) + H^+(aq) + Cl^-(aq)$$

Suggest why tin(II) chloride is usually dissolved in dilute hydrochloric acid.

b. Tin can also exist in the +4 oxidation state. [3]

$$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq) \qquad E^\ominus = +0.15\,V$$

Vanadium can be reduced from an oxidation state of +4 to +3 according to the equation:

$$VO^{2+}(aq) + 2H^+(aq) + e^- \rightleftharpoons V^{3+}(aq) + H_2O(l) \qquad E^\ominus = +0.34\,V$$

(i) Calculate the cell potential, $E^{\ominus}$, and the standard free energy, $\Delta G^{\ominus}$, change for the reaction between the $VO^{2+}$ and $Sn^{2+}$ ions, using sections 1 and 2 of the data booklet.

$E^{\ominus}$:

$\Delta G^{\ominus}$:

(ii) Deduce, giving your reason, whether a reaction between $Sn^{2+}(aq)$ and $VO^{2+}(aq)$ would be spontaneous.

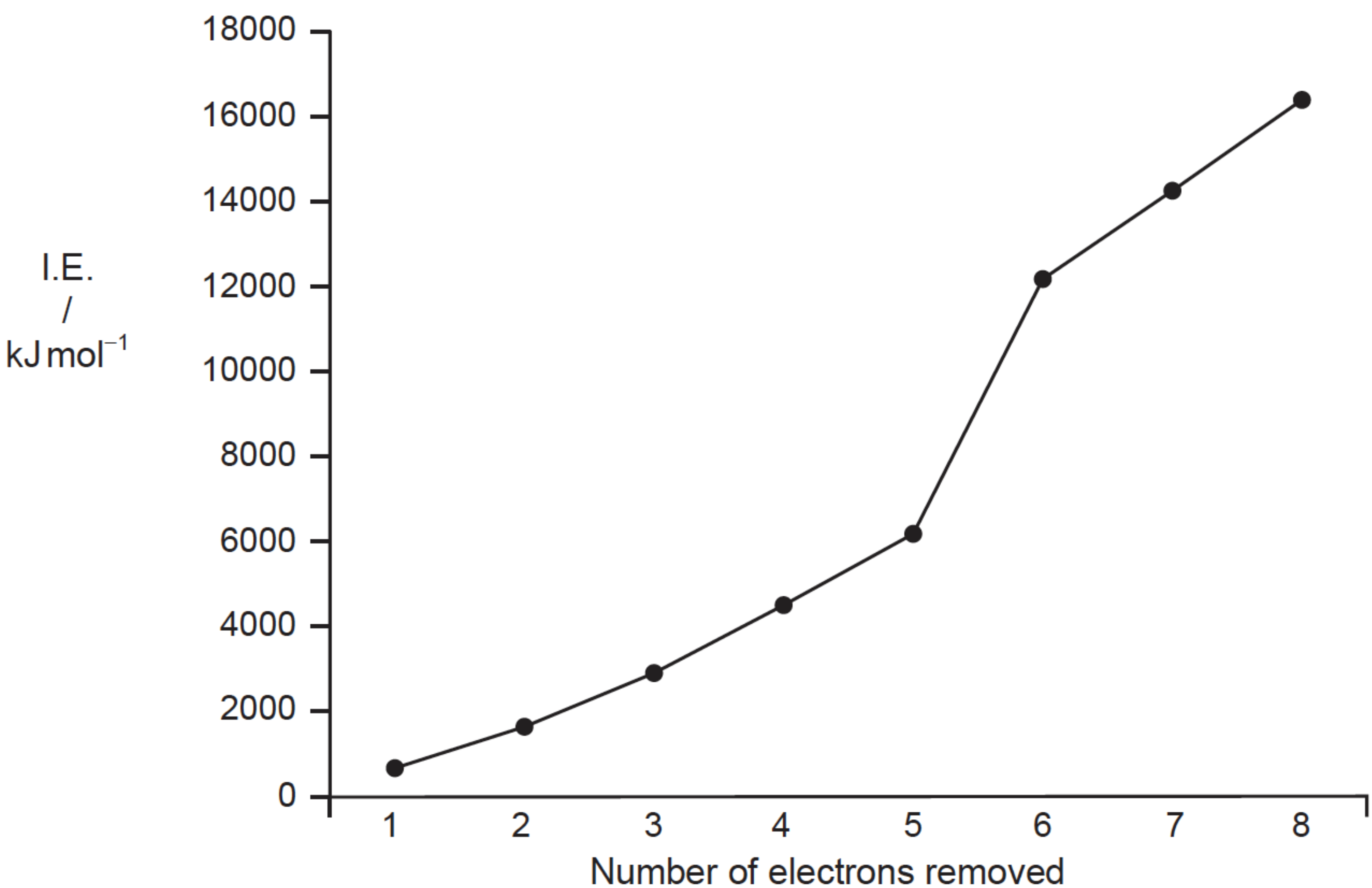
c. Outline, giving the **full** electron configuration of the vanadium atom, what is meant by the term transition metal. [2]

d. In an aqueous solution of vanadium(III) chloride, the vanadium exists as $[V(H_2O)_6]^{3+}$, $[VCl(H_2O)_5]^{2+}$ or $[VCl_2(H_2O)_4]^{+}$ depending on the concentration of chloride ions in the solution. [3]

(i) Describe how $Cl^-$ and $H_2O$ bond to the vanadium ion.

(ii) Outline what would happen to the wavelength at which the vanadium complex ions would absorb light as the water molecules are gradually replaced by chloride ions, using section 15 of the data booklet.

e. Eight successive ionisation energies of vanadium are shown in the graph below: [6]




(i) State the sub-levels from which each of the first four electrons are lost.

First: Second: Third: Fourth:

(ii) Outline why there is an increase in ionization energy from electron 3 to electron 5.

(iii) Explain why there is a large increase in the ionization energy between electrons 5 and 6.

(iv) Vanadium is comprised almost entirely of $^{51}V$. State the number of neutrons an atom of $^{51}V$ has in its nucleus.