PART A
WRITTEN QUESTIONS

QUESTION 1

(a) Write balanced ionic equations for the following chemical reactions. Include states of matter, but do not include spectator ions.

(i) Solid sodium sulfite reacts with hydrogen peroxide solution to give sodium sulfate solution and water.

$$Na_2SO_{3\,(s)} + H_2O_{2\,(aq)} \longrightarrow 2Na^+_{(aq)} + SO_4^{2-}{}_{(aq)} + H_2O_{(l)}$$

(ii) Hydroiodic acid is added to lead(II) nitrate solution giving an orange precipitate of lead(II) iodide.

$$2I^-_{(aq)} + Pb^{2+}_{(aq)} \longrightarrow PbI_{2\,(s)}$$

(iii) Potassium carbonate solution reacts with dilute sulfuric acid to give gaseous carbon dioxide and other products.

$$CO_3^{2-}{}_{(aq)} + 2H^+_{(aq)} \longrightarrow CO_{2\,(g)} + H_2O_{(l)}$$

(b) For the chemical species $^{32}S^{2-}$ provide the following information:

(i) Atomic number: 16

(ii) Mass number: 32

(iii) Number of protons: 16

(iv) Number of neutrons: 16

(v) Number of electrons: 18

(vi) How many atoms are there in total in 1 g of pure water

$$n_{H_2O} = \frac{1\ g}{18.016\ g\ mol^{-1}} = 0.0555\ mol$$

1 mol of H_2O has 2 mol of H atom & 1 mol of O atom

$\therefore$ Total no. of atoms $= 3 \times 0.0555 \times 6.022 \times 10^{23}$

$= 1 \times 10^{23}$ atom

(vii) Write the formula for the conjugate base of the acid $H_2AsO_4^-$:

$HAsO_4^{2-}$

QUESTION 2

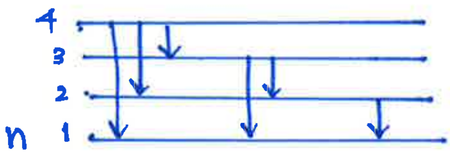
(a) (i) Use the Rydberg equation: $\frac{1}{\lambda} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$ where $R_H = 1.097 \times 10^7 \text{ m}^{-1}$

to calculate the wavelength of light emitted when a hydrogen atom undergoes a transition from the $n = 4$ electronic state to the ground electronic state.

$$\frac{1}{\lambda} = 1.097 \times 10^7 \text{ m}^{-1} \left(\frac{1}{1^2} - \frac{1}{4^2}\right) = 1.028 \times 10^7 \text{ m}$$

$$\therefore \lambda = 9.723 \times 10^{-8} \text{ m} = 97.23 \text{ nm.}$$

(ii) An emission spectrum from hydrogen atoms excited into the $n = 4$ electronic state displays six lines. Write down the six pairs of values for n_1 and n_2 for these lines.



n_2	n_1	n_2	n_1
4 →	3	3 →	2
4 →	2	3 →	1
4 →	1	2 →	1

(iii) Write down the values for n_1 and n_2 for the transition from (ii) above which gives the shortest wavelength of light.

4 → 1 $n_1 = 1$

$n_2 = 4$

(b) Using '$1s^2$…' notation, write the complete ground state electronic configurations of the following gaseous atoms and ions.

(i) Zn $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$

(ii) Cu^{2+} $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9$

(c) Write down all the species from part (b) which are paramagnetic. Cu^{2+}

(d) The ions F^-, Mg^{2+}, O^{2-}, N^{3-} and Na^+ are isoelectronic.

(i) What does isoelectronic mean? Atoms and ions that have same number of electrons are said to be isoelectronic.

(ii) Why do these ions not have identical radii ?

These ions have different no. of protons, so different electrostatic attraction on the electron cloud, which makes the ionic radii different.

QUESTION 3

(a) In the "Chemical Equilibrium" laboratory experiment that you performed this Semester, you investigated the reactions of calcium hydroxide and its products:

Reaction 1. Carbon dioxide gas was bubbled through a saturated solution of calcium hydroxide. A white precipitate immediately formed.

Reaction 2. The mixture obtained in Reaction 1 was treated by passing more carbon dioxide through it. The white precipitate disappeared.

Reaction 3. The solution obtained in Reaction 2 was heated to boiling. The white precipitate reappeared.

Write net ionic chemical equations for each of the chemical reactions described above.

Reaction 1:

$$CO_{2\,(g)} + Ca^{2+}_{(aq)} + 2OH^-_{(aq)} \longrightarrow CaCO_{3\,(s)} + H_2O_{(l)}$$

Reaction 2:

$$CaCO_{3\,(s)} + CO_{2\,(g)} + H_2O_{(l)} \longrightarrow Ca^{2+}_{(aq)} + 2HCO_3^-{}_{(aq)}$$

Reaction 3:

$$Ca^{2+}_{(aq)} + 2HCO_3^-{}_{(aq)} \longrightarrow CO_{2\,(g)} + CaCO_{3\,(s)} + H_2O_{(l)}$$

(b) In the "Chemical Equilibrium" laboratory experiment that you performed this Semester, you investigated the reaction of chromate ions with nitric acid to give dichromate ions. Write a net ionic equation for this reaction.

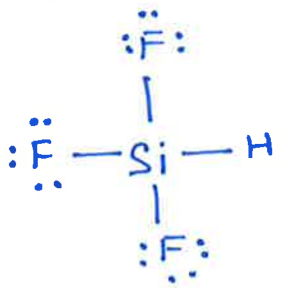
$$2CrO_4^{2-}{}_{(aq)} + 2H^+_{(aq)} \longrightarrow Cr_2O_7^{2-}{}_{(aq)} + H_2O_{(l)}$$

(c) Barium nitrate solution was added to the solution obtained in part (b). No change was observed. However when sodium hydroxide solution was subsequently added, a precipitate formed. Write a net ionic equation for the precipitation reaction.

$$2Ba^{2+}_{(aq)} + Cr_2O_7^{2-}{}_{(aq)} + 2OH^-_{(aq)} \longrightarrow 2BaCrO_{4\,(s)} + H_2O_{(l)}$$

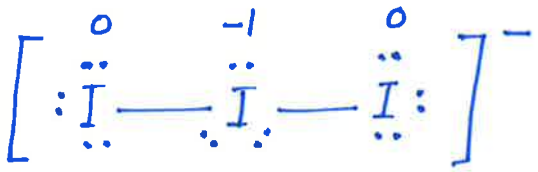
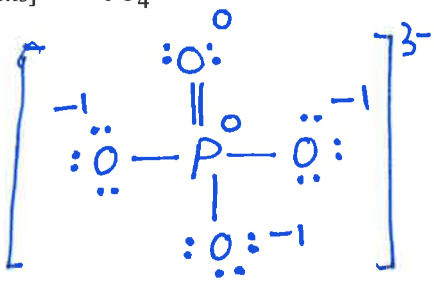
QUESTION 4

(a) For each of the following molecules draw a Lewis diagram, describe the shape of the molecule, describe the hybridisation at the central atom, and state whether the molecule is polar.

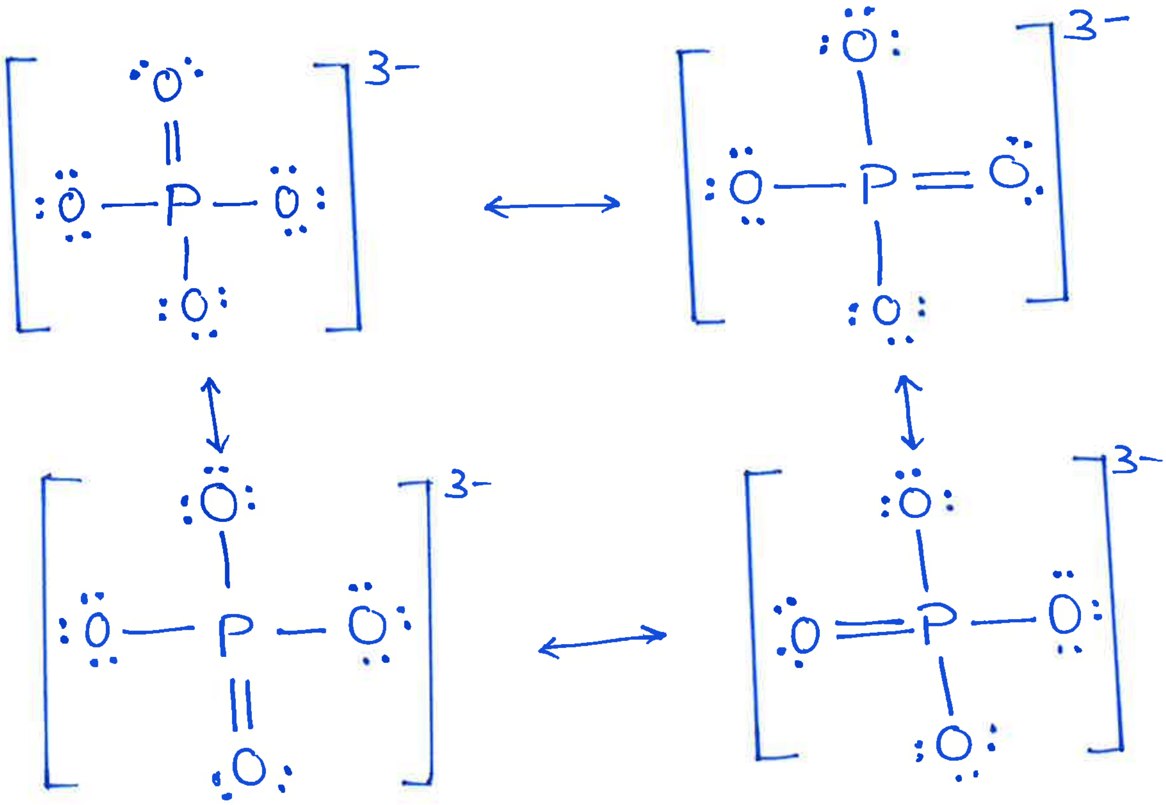
(i) SiF_3H [2 marks]	(ii) OF_2 [2 marks]
	
Shape: [1 mark] tetrahedral	Shape: [1 mark] bent
Hybridisation: [1 mark] sp^3	Hybridisation: [1 mark] sp^3
Polar (yes or no)? [1 mark] yes	Polar (yes or no)? [1 mark] yes

QUESTION 5

(a) For each of the following ions draw a Lewis diagram (with most likely assignments of any formal charges to the appropriate atoms), describe the shape of the ion, and describe the hybridisation at the central atom.

(i) [2 marks] I_3^-	(ii) [2 marks] PO_4^{3-}
	
Shape: [1 mark] linear	Shape: [1 mark] tetrahedral
Hybridisation: [1 mark] sp^3d	Hybridisation: [1 mark] sp^3

(b) Draw four resonance structures for PO_4^{3-}. [2 marks]



QUESTION 6

(a) Indicate the types of intermolecular forces possible between the two molecules listed as pairs in the table below, by writing the word 'yes' or 'no' in the space provided to indicate if each force is possible or not. [5 marks]

The two molecules	dipole–dipole forces	dispersion forces	hydrogen bonding
(i) H_2O and CH_3OH	yes	yes	yes
(ii) N_2 and CN	no	yes	no
(iii) CF_3OCH_3 and H_2O	yes	yes	yes
(iv) CH_2F_2 and CH_2F_2	yes	yes	no

(b) For each of the solutes in Column 1 of the table below, choose the solvent from Column 2 in which it will have the greatest solubility, and write the name of the chosen solvent into Column 3. [2 marks]

Column 1: solute	Column 2: solvent choices	Column 3: chosen solvent
sodium acetate	methanol, dichloromethane, carbon tetrachloride	methanol
methane	methanol, dichloromethane, carbon tetrachloride	carbon tetrachloride

(c) For each molecule in Column 1 of the table below, choose the molecule with the highest boiling point and the molecule with the lowest boiling point. Writh your answer in Column 2 (highest boiling point) and Column 3 (lowest boiling point). [3 marks]

Column 1: Molecule choices	Column 2: Highest boiling point	Column 3: Lowest boiling point
Methane, methanol, propanol, propane	propanol	methane
Hexane, butane, octane, dodecane	dodecane	butane
HF, HCl, HBr, HI	HF	HCl

QUESTION 7

(a) At 448 °C, K_c for the reaction of dihydrogen and diiodine is 50.5:

$H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)$

Use a calculation to predict whether a system containing a 2.00 L vessel at 448 °C contains 20.0 mmol of HI, 10.0 mmol of H_2 and 30.0 mmol of I_2, is at equilibrium? If not, what does your calculation tell you about spontaneous direction of the reaction?

$[HI] = 10.0 \text{ mmol L}^{-1}$

$[I_2] = 15.0 \text{ mmol L}^{-1}$

$[H_2] = 5.0 \text{ mmol L}^{-1}$

$$Q = \frac{[HI]^2}{[H_2][I_2]} = \frac{(10.0)^2}{(5.0)(15.0)} = 1.33$$

$\therefore Q < K_c$ It is not at equilibrium.

The system proceeds from left to right (forward reaction) to reach equilibrium.

(b) Calculate $\Delta H°$ (298 K) for:

$2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$

using these enthalpies of formation (at 298K): $\Delta_f H°$ (ZnS) = –206 kJ mol^{-1}
$\Delta_f H°$ (ZnO) = –350 kJ mol^{-1}
$\Delta_f H°$ (SO_2) = –297 kJ mol^{-1}

$$\Delta H° = \Sigma_i m_i \Delta_f H°_{product\ i} - \Sigma_i m_i \Delta_f H°_{reactant\ i}$$
$$= (2 \times -350 + 2 \times -297) - (2 \times -206)$$
$$= -882 \text{ kJ mol}^{-1}$$

where m_i is the amount (mol)

(c) Solid copper has a heat capacity of 0.39 J K^{-1} g^{-1}. Calculate the heat required to increase the temperature of 1.00 kg of solid copper by 80 °C.

$$q = mc\Delta T$$
$$= 1.00 \times 10^3 \text{ g} \times 0.39 \text{ J K}^{-1}\text{g}^{-1} \times 80 \text{ K}$$
$$= 31 \text{ kJ}$$

(d) 2.315 g of glucose ($C_6H_{12}O_6$, molar mass = 180.2 g mol^{-1}) was burned in a bomb calorimeter to produce $CO_2(g)$ and H_2O (l). The bomb had a heat capacity of 1.800 × 10^3 J K^{-1} and was immersed in 2000 g of water (C_p = 4.184 JK^{-1}g^{-1}). The temperature change of the bomb and surrounding water was +3.52 °C. Calculate the molar internal energy change for the combustion of glucose.

$$n_{C_6H_{12}O_6} = \frac{2.315}{180.2} = 0.01285 \text{ mol}$$

heat absorbed by calorimeter (bomb) $= 1.800 \times 10^3 \times 3.52$
$= 6.34 \times 10^3$ J

heat absorbed by water $= 2000 \times 4.184 \times 3.52 = 2.95 \times 10^4$ J

$\therefore$ Total heat of the reaction, $q_v = -3.58 \times 10^4$ J

For a reaction, carried out in a bomb calorimeter under constant volume condition, q_v = internal energy $= -3.58 \times 10^4$ J

$\therefore$ Molar internal energy $= -3.58 \times 10^4 \text{ J} / 0.01285 \text{ mol} = -2.79 \times 10^3 \text{ kJ mol}^{-1}$

QUESTION 8

(a) Acetic acid can be manufactured by reacting methanol with carbon monoxide:
$CH_3OH(l) + CO(g) \rightarrow CH_3COOH(l)$

Thermodynamic data (at 25 °C, standard state = 1 bar)		
	$\Delta_f H^\circ$ / kJ mol–1	S°/ J K–1 mol–1
$CH_3COOH(l)$	–490	160
$CH_3OH(l)$	–240	130
$CO(g)$	–110	200

(*i*) Calculate ΔH° for the above reaction at 25 °C.

$$\Delta H^0 = \sum_i m_i \Delta_f H^0 \text{product}_i - \sum_i m_i \Delta_f H^0 \text{reactant}_i$$
$$= (-490) - (-240 + -110) = -140 \text{ kJmol}^{-1}$$

where m_i is the amount (mol)

(*ii*) Calculate ΔS° for the above reaction at 25 °C.

$$\Delta S^0 = \sum_i m_i S^0 \text{product}_i - \sum_i m_i S^0 \text{reactant}_i$$
$$= (160) - (130 + 200) = -170 \text{ JK}^{-1}\text{mol}^{-1}$$

(*iii*) Calculate ΔG° for the above reaction at 25 °C.

$$\Delta G^0 = \Delta H^0 - T\Delta S^0$$
$$= -140 \text{ kJmol}^{-1} - 298 \text{K} \times (-170 \times 10^{-3}) \text{kJK}^{-1}\text{mol}^{-1}$$
$$= -89.3 \text{ kJmol}^{-1}$$

(*iv*) Calculate the equilibrium constant for the above reaction at 25 °C.

$$\Delta G^0 = -RT\ln K$$
$$\ln K = -\frac{\Delta G^0}{RT}$$
$$= -\frac{(-89.3 \times 10^3 \text{ J mol}^{-1})}{8.314 \text{ JK}^{-1}\text{mol}^{-1} \times 298\text{K}}$$
$$= 36.04$$
$$K = 4.5 \times 10^{15}$$

(b) Fill in the missing entries in the table below for an aqueous solution at 25 °C.

$[H^+]$ /mol L^{-1}	pH	$[OH^-]$ /mol L^{-1}	pOH
1×10^{-6}	6.0	1×10^{-8}	8.0

(c) Write down the formula and name for the conjugate ~~acid~~ base of each of these species:

	Formula of conjugate base	Name of conjugate base
(i) $HClO_2$	ClO_2^-	chlorite ion
(ii) HSO_3^-	SO_3^{2-}	sulfite ion

(d) Use the data in the table to *circle* the correct answer to the following questions.

(*i*) The salt which gives more acidic solution:

$AlCl_3$ or CH_3COONa

(*ii*) The stronger base: NH_3 or CH_3COO^-

	pK_a at 25 °C
Al^{3+}	4.96
CH_3COOH	4.76
NH_4^+	9.24

QUESTION 9

(a) Benzoic acid, a monoprotic acid with the formula C_6H_5COOH (HA), has a pK_a value of 4.20.

(*i*) Calculate the pH of 0.103 M benzoic acid.

$$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons A^-_{(aq)} + H_3O^+_{(aq)}$$

	HA		A^-	H_3O^+
I	0.103		0	0
Δ	$-x$		$+x$	$+x$
F	$0.103-x$		$+x$	$+x$

$$K_a = \frac{x^2}{0.103-x}$$

$$10^{-4.20} = \frac{x^2}{0.103}$$ assume $x < .103$

$$x = 2.55 \times 10^{-3}$$

$pH = 2.60$

(*ii*) Calculate the pH after adding 40.0 mL of 0.206 M NaOH to 80.0 mL of 0.103 M benzoic acid.

$n_{HA} = 8.24 \times 10^{-3}$ mol

$n_{OH^-} = 8.24 \times 10^{-3}$ mol

At equivalence pt.

n_{A^-} produced $= 8.24 \times 10^{-3}$ mol

$\therefore [A^-] = 0.0687$ mol L^{-1}

$$A^-_{aq} + H_2O_{(l)} \rightleftharpoons HA_{aq} + OH^-_{aq}$$

$0.0687-x \qquad x \qquad x$

$$K_b = \frac{x^2}{0.0687-x}$$ assume $x < .0687$

$$10^{-9.80} = \frac{x^2}{.0687}$$

$x = 3.3 \times 10^{-6}$

$pOH = 5.48$

$pH = 8.52$

(b) What is the oxidation number of the element indicated in each of the following species?

Species	S_8	H_3PO_4	Cl_2O_7	HCO_3^-
Element	S	P	Cl	C
Oxidation Number	0	+5	+7	+4

(c) Hydrogen peroxide (H_2O_2) is oxidised by permanganate ion (MnO_4^-). Write balanced half equations and the overall chemical equation for the reaction of aqueous hydrogen peroxide with permanganate ions (MnO_4^-) in acid solution to produce a solution of manganese(II) ions.

$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O \;] \times 2$$

$$H_2O_2 \longrightarrow 2H^+ + O_2 + 2e^- \;] \times 5$$

$$2MnO_4^-{}_{(aq)} + 5H_2O_{2(aq)} + 6H^+_{(aq)} \longrightarrow 2Mn^{2+}_{(aq)} + 5O_{2(g)} + 8H_2O_{(l)}$$

QUESTION 10

(a) For the cell represented by this diagram: Cd(s) | Cd^{2+}(aq) || Ag^{+}(aq) | Ag(s),

(*i*) Write a balanced half-cell equation for the reaction at the anode.

$$Cd_{(s)} \longrightarrow Cd^{2+}_{(aq)} + 2e$$

(*ii*) Write a balanced equation showing the overall reaction occuring in the cell.

$$Cd_{(s)} + 2Ag^{+}_{(aq)} \longrightarrow Cd^{2+}_{(aq)} + 2Ag_{(s)}$$

(*iii*) Given the standard reduction potentials: $E°$ (Cd^{2+} | Cd) = –0.40 V and $E°$ (Ag^{+} | Ag) = +0.80 V, calculate the standard cell potential.

$$E^{o}_{cell} = E^{o}_{cathode} - E^{o}_{anode} = (0.80) - (-0.40)$$
$$= 1.20\ V$$

(*iv*) Calculate $\Delta G°$ for the cell.

$$\Delta G^{o} = -nFE^{o}_{cell} = -2 \times 96485\ Cmol^{-1} \times 1.20V$$
$$= -232\ kJmol^{-1}$$

(b) List two reasons (maximum 3 words each) why lithium metal is often used in batteries for portable electronic devices.

Reason 1: Light weight metal

Reason 2: has most negative reduction potential

(c) Briefly explain why cracks develop in steel reinforced concrete when measures are not taken to prevent corrosion of the reinforcing rods.

Iron oxide (rust) takes up more volume than iron metal. Expansion causes cracks.

(d) Calculate the mass of zinc produced in 8.00 min by the electrolysis of molten $ZnCl_2$ using a current of 300 A.

$$Q = I \times t$$
$$= 300A \times (8.00 \times 60)s$$
$$= 1.44 \times 10^{5}\ C$$
$$\therefore n_e = Q/F = \frac{1.44 \times 10^{5} C}{96485\ C\ mol^{-1}}$$
$$= 1.49\ mol\ of\ e^{-}$$

$$Zn^{2+}_{(aq)} + 2e \rightarrow Zn_{(s)}$$
$$\therefore n_{Zn} = \frac{1}{2} \times 1.49$$
$$= 0.745$$
$$mass_{Zn} = 0.745 \times 65.39$$
$$= 48.8\ g$$

PART B
MULTIPLE CHOICE

There are 10 multiple choice questions in this section, each worth 2 marks. Each multiple choice consists of a statement or question followed by 5 possible choices. Select the choice that best answers the statement or question by CIRCLING the appropriate letter.

There is only one correct answer for each question.

- TRANSFER YOUR ANSWERS TO THE GENERALISED ANSWER SHEET WITH PENCIL.
- YOUR MARK FOR THIS PART WILL BE DETERMINED FROM YOUR ENTRIES ON THE GENERALISED ANSWER SHEET.

THIS SECTION OF THE PAPER IS NOT AVAILABLE

DATA SHEET

0 °C = 273 K

1 atm = 760 mmHg = 101.3 kPa = 760 Torr

Ideal Gas Constant $R = 8.314\ \text{J mol}^{-1}\ \text{K}^{-1} = 0.08206\ \text{L atm mol}^{-1}\ \text{K}^{-1}$

Avogadro Number $N_A = 6.022 \times 10^{23}\ \text{mol}^{-1}$

1 atm = 760 mmHg = 101.3 kPa = 760 Torr = 1.013 bar

Faraday Constant $F = 96{,}485\ \text{C mol}^{-1}$

Nernst Equation $E_{cell} = E°_{cell} - \frac{RT}{nF} \ln Q$

Faraday Equation $Q = i \times t$ = amount electrons (mole) $\times F$

Planck Constant $h = 6.626 \times 10^{-34}\ \text{J s}$

Speed of Light $c = 2.998 \times 10^{8}\ \text{m s}^{-1}$

Planck Equation $E = h\nu$

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Periodic Group Numbers

The Periodic Table

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.003
3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.88	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 ^{99}**Tc** (98.91)	44 **Ru** 101.1	45 **Rh** 102.9	46 **Pd** 106.4	47 **Ag** 107.9	48 **Cd** 112.4	49 **In** 114.8	50 **Sn** 118.7	51 **Sb** 121.8	52 **Te** 127.6	53 **I** 126.9	54 **Xe** 131.3
55 **Cs** 132.9	56 **Ba** 137.3	57 **La** 138.9	72 **Hf** 178.5	73 **Ta** 180.9	74 **W** 183.9	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.1	79 **Au** 197.0	80 **Hg** 200.6	81 **Tl** 204.4	82 **Pb** 207.2	83 **Bi** 209.0	84 ^{210}**Po** (210.0)	85 ^{210}**At** (210.0)	86 ^{222}**Rn** (222.0)
87 ^{223}**Fr** (223.0)	88 ^{226}**Ra** (226.0)	89 ^{227}**Ac** (227.0)															

58 **Ce** 140.1	59 **Pr** 140.9	60 **Nd** 144.2	61 ^{145}**Pm** (144.9)	62 **Sm** 150.4	63 **Eu** 152.0	64 **Gd** 157.3	65 **Tb** 158.9	66 **Dy** 162.5	67 **Ho** 164.9	68 **Er** 167.3	69 **Tm** 168.9	70 **Yb** 173.0	71 **Lu** 175.0
90 **Th** 232.0	91 ^{231}**Pa** (231.0)	92 **U** 238.0	93 ^{237}**Np** (237.0)	94 ^{239}**Pu** (239.1)	95 ^{243}**Am** (243.1)	96 ^{247}**Cm** (247.1)	97 ^{247}**Bk** (247.1)	98 ^{252}**Cf** (252.1)	99 ^{252}**Es** (252.1)	100 ^{257}**Fm** (257.1)	101 ^{256}**Md** (256.1)	102 ^{259}**No** (259.1)	103 ^{260}**Lr** (260.1)

KEY

Atomic N° → 6
Symbol → **C**
Atomic Weight → 12.01

( ) is the relative atomic mass of the most common radioactive isotope, the mass number of which is given as a superscript.

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