Topics in the June 2006 Exam Paper for CHEM1001

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- Stoichiometry
- Atomic Energy Levels

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- Molecules and Ions
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- Gas Laws

	Marks
• Balance the following nuclear reactions by identifying the missing nuclear particle.	2

$$^{234}_{90}\text{Th} \rightarrow \boxed{^{234}_{91}\textbf{Pa}} + {}^{0}_{-1}\text{e}$$

$$^{234}_{92}\text{U} \rightarrow \boxed{^{230}_{90}\textbf{Th}} + {}^{4}_{2}\text{He}$$

• A nugget contains 2.6×10^{24} atoms of gold. What amount of gold (in mol) is in this nugget and what is its mass (in kg)? **2**

One mole of gold corresponds to Avogadro's number, 6.022×10^{23}, atoms. 2.6×10^{24} atoms therefore corresponds to:

$$\textbf{number of moles} = \frac{\textbf{number of atoms}}{\textbf{Avogadro's number}} = \frac{2.6 \times 10^{24}}{6.022 \times 10^{23}} = 4.3\ \textbf{mol}$$

As one mole of gold has a mass, corresponding to the atomic mass, of 196.97 g. 4.3 mol of gold therefore corresponds to:

mass = number of moles × atomic mass = 4.3 × 196.97 = 850 g = 0.85 kg

(Note that the number of atoms is given to 2 significant figures in the question and this is reflected in the answers).

Amount: **4.3 mol**	Mass: **0.85 kg**

• What element has the ground state electronic arrangement of $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$? **1**

Phosphorus

• A mobile phone sends signals at about 850 MHz (1 MHz = 1×10^6 Hz). What is the wavelength of this radiation? **2**

The frequency, ν =850 MHz = 850×10^6 Hz, is related to the wavelength, λ, by the equation:

$$c = \lambda\nu \ \text{ or } \ \lambda = \frac{c}{\nu} \ \text{ where c is the speed of light.}$$

$$\textbf{Therefore, wavelength} = \lambda = \frac{2.998 \times 10^{8}\ \text{m s}^{-1}}{850 \times 10^{6}\ \text{m}} = 0.35\ \text{m}$$

Wavelength = **0.35 m**

Marks

- Account for why solid metals can conduct an electric current, but solid ionic compounds cannot. **3**

The crystal structure of a metal consists of a lattice of positively charged nuclei surrounded by a "sea of electrons". These electrons are free to move under the influence of an electric field so can conduct the current.

An ionic solid consists of a lattice of positive and negative ions, packed together to minimise repulsion and maximise attraction. The atomic nuclei are fixed in place and all the electrons are localised around them so they are unable to conduct the current (They can conduct current when molten as the ions are then free to move.)

- Complete the entries in the following table. **3**

Element name	Symbol	Mass number	Atomic number	Number of electrons	Number of neutrons	$^{m}_{z}X$
lithium	**Li**	7	3	**3**	**4**	$^{7}_{3}\mathbf{Li}$
copper	Cu	**64**	**29**	29	**35**	$^{64}_{29}\mathbf{Cu}$
aluminium	**Al**	**27**	13	**13**	14	$^{27}_{13}\mathbf{Al}$

- Give the formula and name of the binary compound formed from the following elements. **2**

	Formula	Name
lithium and oxygen	$\mathbf{Li_2O}$	**lithium oxide**
calcium and hydrogen	$\mathbf{CaH_2}$	**calcium hydride**

Marks 7

- The complete combustion of propane, C_3H_8, in air gives water and carbon dioxide as the products? Write a balanced equation for this reaction.

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$

What mass of oxygen is required for the complete combustion of 454 g of propane and what masses of CO_2 and H_2O are produced?

The molar mass of propane is (3 × 12.01 (C)) + (8 × 1.008 (H)) = 44.094. Therefore, 454 g corresponds to:

$$\textbf{number of moles} = \frac{\textbf{mass}}{\textbf{molar mass}} = \frac{454}{44.094} = 10.3 \textbf{ mol}$$

5 mol of $O_2(g)$ is required for every 1 mol of propane. Therefore, 5 × 10.3 = 51.5 mol of O_2 is required. The molar mass of O_2 is (2 × 16.00) = 32.00 so the mass of O_2 required is:

mass = number of moles × molar mass = 51.5 × 32.00 = 1650 g = 1.65 kg

3 mol of CO_2 and 4 mol of H_2O are produced for every 1 mol of propane. Therefore, 3 × 10.3 = 30.9 mol of CO_2 and 4 × 10.3 = 41.2 mol of H_2O are produced.

The molar mass of CO_2 is (12.01 (C)) + (2 × 16.00 (O)) = 44.01 and the molar mass of H_2O is (2 × 1.008 (H)) + (16.00 (O)) = 18.016.

The masses of CO_2 and H_2O produced are therefore:

mass of CO_2 = 30.9 × 44.01 = 1360 g = 1.36 kg
mass of H_2O = 41.2 × 18.016 = 742 g = 0.742 kg

(Note that the mass of propane is given to three significant figures in the question and this is reflected in each answer).

Explain the "law of conservation of mass". Show whether or not the above combustion conforms to this law.

The law of conservation of mass states that mass may neither be created nor destroyed. In this reaction:

mass of reactants = 0.454 kg (C_3H_8) + 1.65 kg (O_2) = 2.10 kg
mass of products = 1.36 kg (CO_2) + 0.742 kg (H_2O) = 2.10 kg

This combustion obeys the law (to 3 significant figures).

Marks 3

- The reaction of methane and water is one way to prepare hydrogen for use as a fuel.

$$CH_4(g) + H_2O(g) \rightarrow CO(g) + 3H_2(g)$$

Which compound is the limiting reactant if you begin with 995 g of methane and 2510 g of water?

The molar mass of methane, CH_4, is (12.01 (C)) + (4 × 1.008 (H)) = 16.042. The number of moles of methane is therefore:

$$\textbf{moles of methane} = \frac{\textbf{mass}}{\textbf{molar mass}} = \frac{\textbf{995}}{\textbf{16.042}} = \textbf{62.0 mol}$$

The molar mass of water, H_2O, is (2 × 1.008 (H)) + (16.00 (H)) = 18.016. The number of moles of water is therefore:

$$\textbf{moles of methane} = \frac{\textbf{mass}}{\textbf{molar mass}} = \frac{\textbf{2510}}{\textbf{18.016}} = \textbf{139 mol}$$

As the reaction is a 1:1 reaction of methane and water, methane is the limiting reagent.

Answer: **methane, CH_4**

What mass of the excess reactant remains when the reaction is completed?

As the reaction is a 1:1 reaction, (139 – 62.0) = 77 mol of H_2O will be left unreacted. This corresponds to a mass of:

mass of water = moles of water × molar mass = 77 × 18.016 = 1400 g = 1.4 kg.

Answer: **1.4 kg**

Marks

- An unknown compound contains carbon and hydrogen only. If 0.0956 g of the compound is burned in oxygen, 0.300 g of CO_2 and 0.123 g of H_2O are isolated. What is the unknown compound's empirical formula?

4

The molar mass of CO_2 is (12.01 (C)) + (2 × 16.00 (O)) = 44.01. The molar mass of H_2O is (2 × 1.008 (H)) + (16.00 (O)) = 18.016. The number of moles of each after burning is therefore:

$$n_{CO_2} = \frac{\text{mass}}{\text{molar mass}} = \frac{0.300}{44.01} = 0.00682\,\text{mol}, \quad n_{H_2O} = \frac{0.123}{18.016} = 0.00683\,\text{mol}$$

The moles of C in the compound is equal to the number of moles of CO_2, as the latter possesses one carbon atom per molecular unit.

The moles of H in the compound is equal to 2 × number of moles of H_2O, as the latter contains two hydrogen atoms per molecular unit.

The C:H ratio is therefore 1:2

Answer: **CH_2**

If its molar mass is found to be 70.1 g mol^{-1}, what is its molecular formula?

If the molar mass = 70.1, the number of moles in 0.0956 g is:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{0.0956}{70.1} = 0.00136\,\text{mol}$$

As 0.00136 mol contains 0.683 mol of carbon, 1 mol contains $\frac{0.00683}{0.00136} = 5.01\,\text{mol}$

As C:H is 1:2, the compound must contain 10 hydrogen atoms.

Answer: **C_5H_{10}**

- What amount (in mol) of chloride ion is contained in 100 mL of 0.25 M magnesium chloride solution?

1

Magnesium chloride dissolves according to the equation, $MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$ so that two moles of chloride is produced for every mole of $MgCl_2$ present. The number of moles of $MgCl_2$ present is :

$$\text{number of moles} = \text{concentration} \times \text{volume} = 0.25 \times \frac{100}{1000} = 0.025\ \text{mol}$$

The number of moles of $Cl^-(aq)$ is therefore 2 × 0.025 = 0.050 mol

Answer: **0.050 mol**

1

- If 25.0 mL of 1.50 M hydrochloric acid is diluted to 500 mL, what is the molar concentration of the diluted acid?

The number of moles of HCl present in 25.0 mL of a 1.50 M solution is:

$$\textbf{number of moles = concentration} \times \textbf{volume} = \mathbf{1.50} \times \frac{\mathbf{25}}{\mathbf{1000}} = \mathbf{0.0375\ mol}$$

This number of moles in a 500 mL solution gives a concentration of:

$$\textbf{concentration} = \frac{\textbf{number of moles}}{\textbf{volume}} = \frac{\mathbf{0.0375}}{\mathbf{(500/1000)}} = \mathbf{0.0750\,M}$$

Answer: **0.0750 M**

Marks

- A 1.00 g sample of ammonium nitrate, NH_4NO_3, is decomposed in a bomb calorimeter causing the temperature of the calorimeter to increase by 6.12 K. The heat capacity of the system is 1.23 kJ °C^{-1}. **3**

Describe this process as either endothermic or exothermic. **exothermic**

What is the molar heat of decomposition for ammonium nitrate?

The heat change is given by q = C × ΔT = 1.23 × 6.12 = 7.53 kJ. As the reaction is exothermic, the heat of decomposition for 1.00 g of NH_4NO_3 is ΔH = -7.53 kJ.

The molar mass of NH_4NO_3 is:

molar mass = (2 × 14.01 (N)) + (4 × 1.008 (H)) + (3 × 16.00 (O)) = 80.052

1.000 g therefore corresponds to $\frac{1.000}{80.052}$ = 0.0125 mol. The molar heat of decomposition is then: $\Delta H = \frac{-7.53}{0.0125} = -602\text{ kJ mol}^{-1}$

Answer: **-602 kJ mol^{-1}**

- Heating $SbCl_5$ causes it to decompose according to the following equation. **4**

$$SbCl_5(g) \rightleftharpoons SbCl_3(g) + Cl_2(g)$$

A sample of 0.50 mol of $SbCl_5$ is placed in a 1.0 L flask and heated to 450 ºC. When the system reaches equilibrium there is 0.10 mol of Cl_2 present. Calculate the value of the equilibrium constant, K_c, at 450 °C.

One mole of Cl_2 is generated by the decomposition of one mole of $SbCl_5$. As 0.10 mol of Cl_2 is present at equilibrium, (0.50 – 0.10) = 0.40 mol of $SbCl_5$ must be left.

One mole of $SbCl_3$ is generated alongside the production of one mole of Cl_2 so the number of moles of $SbCl_3$ = number of moles of Cl_2 = 0.10 mol.

The volume of the flask is 1.0 L so the concentration = $\frac{\text{number of moles}}{\text{volume}}$. The concentrations are therefore: [$SbCl_5$(g)] = 0.40 M, [Cl_2(g)] = [$SbCl_3$(g)] = 0.10 M.

The equilibrium constant in terms of concentrations, K_c, is therefore:

$$K_c = \frac{[Cl_2(g)][[SbCl_3(g)]}{[SbCl_5(g)]} = \frac{(0.10)\times(0.10)}{(0.40)} = 0.025$$

Answer: **K_c = 0.025**

Marks

- Consider a cell composed of the following half-reactions. **3**

$$Ce^{4+}(aq) + e^{-} \rightarrow Ce^{3+}(aq)$$

$$Cr(s) \rightarrow Cr^{3+}(aq) + 3e^{-}$$

What is the balanced equation for the spontaneous reaction?

$3Ce^{4+}(aq) + Cr(s) \rightarrow 3Ce^{3+}(aq) + Cr^{3+}(aq)$

What is the value of $E°$ for the cell? Relevant standard reduction potentials are on the data sheet.

The electrode potentials are:

$Ce^{4+}(aq) + e^{-} \rightarrow Ce^{3+}(aq)$ E° = +1.72 V
$Cr(s) \rightarrow Cr^{3+}(aq) + 3e^{-}$ E° = +0.74 V (reversed as oxidation required)

The standard cell potential is therefore: E° = (+1.72) + (+0.74) = +2.46 V

Answer: +**2.46 V**

- What does the superscript "o" mean in the symbol $\Delta H_f°$? **1**

The enthalpy change corresponds to all reactants and products being in their standard states (gases at pressures of 100 kPa, solutions of 1 M concentration and elements in their common form at 100 kPa and 273 K)

Marks

- Calculate the standard heat of reaction for the following reaction. **2**

$$Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$$

Data: $\Delta H^{o}{}_{f} = +64.4 \text{ kJ mol}^{-1}$ for $Cu^{2+}(aq)$

$\Delta H^{o}{}_{f} = -152.4 \text{ kJ mol}^{-1}$ for $Zn^{2+}(aq)$

The enthalpy of the reaction is given by:

$$\begin{aligned} \Delta_{rxn}H^{o} &= \sum m\Delta_f H^{o}(\text{products}) - \sum n\Delta_f H^{o}(\text{reactants}) \\ &= [\Delta_f H^{o}(Zn^{2+}(aq))] - [\Delta_f H^{o}(Cu^{2+}(aq))] \\ &= (-152.4) - (+64.4) = -216.8 \text{ kJ mol}^{-1} \end{aligned}$$

(Note that $\Delta_f H^{o}(Zn(s))$ and $\Delta_f H^{o}(Cu(s))$ are both zero as these elements are in the standard states).

Answer: **-216.8 kJ mol^{-1}**

- Write a balanced **ionic** equation for the reaction of solid sodium hydrogencarbonate, $NaHCO_3$, and dilute sulfuric acid, H_2SO_4. **2**

$$\mathbf{NaHCO_3(s) + H^{+}(aq) \rightarrow Na^{+}(aq) + H_2O(l) + CO_2(g)}$$

Marks

- Calculate the mass of silver nitrate, $AgNO_3$, required to make 500 mL of 0.200 M aqueous solution. **4**

The number of moles of $AgNO_3$ in 500 mL of a 0.200 M solution is:

$$\textbf{number of moles = concentration} \times \textbf{volume} = 0.200 \times \frac{500}{1000} = 0.100 \textbf{ mol}$$

The formula mass of $AgNO_3$ is:

formula mass = (107.87 (Ag)) + (14.01 (N)) + (3 × 16.00 (O)) = 169.88

The mass of 0.100 mol is therefore 0.100 × 169.88 = 17.0 g

Answer: **17.0 g**

Calculate the time required (in minutes) to deposit 7.0 g of silver from a 0.200 M silver nitrate solution using a current of 4.5 A.

The number of moles of silver in 7.0 g is $\frac{\textbf{mass}}{\textbf{atomic mass}} = \frac{7.0}{107.87} = 0.065 \textbf{ mol}$

To deposit silver requires reduction of Ag^+(aq), requiring 1 mole of electrons per mole of silver. Hence, 0.065 mol of electrons is required. This corresponds to a charge of Q = nF = 0.065 × F = 0.065 × 96485 = 6300 C.

As Q = I × t, the time taken to deliver this charge at a current of 4.5 A is:

$$t = \frac{Q}{I} = \frac{6300}{4.5} = 1400 \text{ s} = 23 \textbf{ minutes}$$

Answer: **23 minutes**

- A lead-acid battery has the following shorthand notation: **4**

  $Pb(s), PbSO_4(s) \mid H^+(aq), SO_4^{2-}(aq) \parallel H^+(aq), SO_4^{2-}(aq) \mid PbO_2(s), PbSO_4(s)$

Which component of the battery is the anode? **$Pb(s), PbSO_4(s)$**

Give the balanced half equation of the reaction that takes place at the anode.

$Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2e^-$

Which component of the battery is the cathode? **$PbO_2(s), PbSO_4(s)$**

Give the balanced half equation of the reaction that takes place at the cathode.

$PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$

Marks
4

- When "dry ice", solid carbon dioxide, is heated to 400 K it becomes gaseous. An 88.0 g sample of solid carbon dioxide is placed into a sealed 100 L container that is initially at a pressure of 1.00 atm and a temperature of 298 K. The container is heated to 400 K. What will be the final pressure inside the container?

The molar mass of CO_2 is (12.01 (C)) + (2 × 16.00 (O)) = 44.01. The sample of 88.0 g therefore corresponds to:

$$\text{moles of } CO_2 = \frac{\text{mass}}{\text{molar mass}} = \frac{88.0}{44.01} = 2.00 \text{ mol}$$

The pressure due to this amount of CO_2 in a container of volume 100 L at 400 K is given by the ideal gas law, PV = nRT, as:

$$p_{CO_2} = \frac{nRT}{V} = \frac{(2.00)\times(0.08206)\times(400)}{100} = 0.656 \text{ atm}$$

The container initially contained a gas with a pressure of 1.00 atm at 298 K. The ideal gas equation can be applied to this gas at the two temperatures and as the number of moles and the volume is constant, this gas will have a pressure at 400 K corresponding to:

$$\frac{p(T_2)}{p(T_1)} = \frac{T_2}{T_1} \text{ or } p(400 \text{ K}) = p(298 \text{ K})\times\frac{400}{298} = 1.00\times\frac{400}{298} = 1.34 \text{ atm}$$

The total pressure at 400 K is therefore 0.656 + 1.34 = 2.00 atm

Answer: **2.00 atm**