Topics in the June 2007 Exam Paper for CHEM1001

Click on the links for resources on each topic.

2007-J-2:

- Elements and Atoms
- Chemical Equations
- Stoichiometry

2007-J-3:

- Lewis Model of Bonding
- VSEPR
- Elements and Atoms

2007-J-4:

- Stoichiometry
- Chemical Equations

2007-J-5:

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2007-J-6:

- Thermochemistry
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2007-J-7:

- Introduction to Electrochemistry
- Electrochemistry
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2007-J-8:

- Electrolytic Cells
- Introduction to Electrochemistry
- Electrochemistry
- Batteries and Corrosion

2007-J-9:

- Chemical Equilibrium

2007-J-10:

- Gas Laws

Marks

- Write balanced equations for the following nuclear reactions. **4**

Naturally occurring thorium 232 undergoes alpha decay.	$^{232}_{90}\mathrm{Th} \rightarrow\ ^{228}_{88}\mathrm{Ra} +\ ^{4}_{2}\mathrm{He}$
A nuclide undergoes beta decay and produces caesium 133.	$^{133}_{54}\mathrm{Xe} \rightarrow\ ^{133}_{55}\mathrm{Cs} +\ ^{0}_{-1}\beta$

- A cook uses a microwave oven to heat up a meal. The wavelength of the radiation is 0.012 m. Calculate the frequency and energy of a photon of this radiation. **2**

The wavelength, λ, is related to the energy and the frequency, ν, by the equations:

$$\mathbf{E = h\nu = \frac{hc}{\lambda} \text{ and } \nu = \frac{c}{\lambda}}$$

Therefore with λ = 0.012 m:

$$\nu = \frac{(2.998\times10^{8})}{(0.012)} = 2.5\times10^{10}\ \mathrm{s^{-1}}$$

$$E = \frac{(6.626\times10^{-34})\times(2.998\times10^{8})}{(0.012)} = 1.7\times10^{-23}\ \mathrm{J}$$

(As the wavelength is given to two significant figures, this limits the accuracy of the answers to also being two significant figures).

Frequency: **2.5×10^{10} s^{-1}**	Energy: **1.7×10^{-23} J**

ANSWER CONTINUES ON THE NEXT PAGE

- What mass of calcium chloride is required to make 250 mL of a 0.1 M solution? **3**

The formula mass of calcium chloride, $CaCl_2$, is

formula mass = 40.08 (Ca) + 2 × 35.45 (Cl) = 110.98

The number of moles in the solution is given by:

$$\textbf{number of moles} = \textbf{concentration} \times \textbf{volume} = 0.1 \times \frac{250}{1000} = 0.025 \textbf{ mol}$$

The mass required is therefore:

mass = number of moles × formula mass = (0.025) × (110.98) = 3 g

Answer: **3 g**

What amount of chloride ions (in mol) is present in 30.0 mL of this solution?

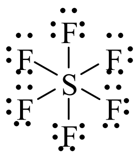
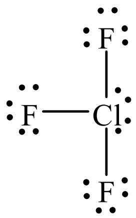
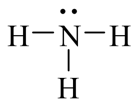
One moles of $CaCl_2(s)$ dissolves to give two moles of $Cl^-(aq)$ ions. Therefore, the number of moles present is:

$$\textbf{number of moles} = \textbf{concentration} \times \textbf{volume} = (2 \times 0.1) \times \frac{30}{1000} = 0.006 \textbf{ mol}$$

Answer: **0.006 mol**

Marks
9

- Complete the following table.

Molecular formula	SF_6	$\mathbf{ClF_3}$	NH_3
Name		chlorine trifluoride	
Lewis structure			
Number of bonding electron pairs on central atom	**6**	**3**	**3**
Number of non-bonding electron pairs on central atom	**0**	**2**	**1**
Molecular shape	**octahedral**	**T-shaped**	**trigonal pyramidal**

2

- Silicon is essential to the computer industry as a major component of chips. It has three naturally occurring isotopes, the relative abundance of each being given below. Calculate the atomic mass of silicon.

Isotope	Mass of isotope (a.m.u.)	Relative abundance
^{28}Si	27.9769	92.23%
^{29}Si	28.9765	4.67%
^{30}Si	29.9738	3.10%

The relative atomic mass of silicon is the weighted average of the masses of its isotopes:

$$\textbf{atomic mass} = \left(27.9769 \times \frac{92.23}{100}\right) + \left(28.9765 \times \frac{4.67}{100}\right) + \left(29.9738 \times \frac{3.10}{100}\right)$$

$$= \mathbf{28.09}$$

(The relative abundances are given to 4 significant figures and limit the accuracy of the answer.)

Answer: **28.09**

Marks 4

- The complete combustion of butane, C_4H_{10}, in air gives water and carbon dioxide as the products. Write a balanced equation for this reaction.

$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$ or

$C_4H_{10}(g) + {}^{13}/_2O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$

What mass of oxygen is required for the complete combustion of 454 g of butane and what masses of carbon dioxide and water are produced?

The molar mass of butane is (4 × 12.01 (C)) + (10 × 1.008 (H)) = 58.12. Therefore, the amount of butane in 454 g is:

$$\text{number of moles of } C_4H_{10} = \frac{\text{mass}}{\text{molar mass}} = \frac{454}{58.12} = 7.81\,\text{mol}$$

From the chemical equation, each mole of C_4H_{10} requires ${}^{13}/_2$ moles of O_2 and produces 4 moles of CO_2 and 5 moles of H_2O.

Therefore:

number of moles of O_2 = ${}^{13}/_2$ × 7.81 = 50.8 mol
number of moles of CO_2 = 4 × 7.81 = 31.2 mol
number of moles of O_2 = 5 × 7.81 = 39.1 mol

The molar masses of O_2, CO_2 and H_2O are:

molar mass of O_2 = 2 × 16.00 = 32.00
molar mass of CO_2 = 12.01 (C) + (2 × 16.00) = 44.01
molar mass of H_2O = (2 × 1.008 (H)) + 16.00 (O)= 18.016

Therefore:

mass of O_2 = number of moles × molar mass = 50.8 × 32.00 = 1620 g = 1.62 kg
mass of CO_2 = 31.2 × 44.01 = 1380 g = 1.38 kg
mass of H_2O = 39.1 × 18.016 = 704 g = 0.704 kg

ANSWER CONTINUES ON THE NEXT PAGE

- During physical activity, lactic acid forms in the muscle tissue and is responsible for muscle soreness. Elemental analysis shows that it contains by mass 40.0% C, 6.71% H and 53.3% O. Determine the empirical formula of lactic acid.

4

	C	H	O
amount in 100 g	**40.0**	**6.71**	**53.3**
ratio (divide by atomic mass)	$\frac{40.0}{12.01} = 3.33$	$\frac{6.71}{1.008} = 6.66$	$\frac{53.3}{16.00} = 3.33$
divide by smallest	$\frac{3.33}{3.33} = 1.00 \sim 1$	$\frac{6.66}{3.33} = 2.00 \sim 2$	$\frac{3.33}{3.33} = 1.00$

The simplest possible ratio of C:H:O is thus 1:2:1 and the empirical formula is CH_2O.

Answer: **CH_2O**

Given that lactic acid has a molar mass of 90.08 g mol^{-1}, determine its molecular formula.

The molecular formula is $(CH_2O)_n$ so the molar mass is:

$$\text{molar mass} = n \times (12.01\ (C) + 2 \times 1.008\ (H) + 16.00\ (O))$$
$$= 30.026n = 90.08 \text{ so } n = 3$$

The molecular formula is thus $(CH_2O)_3$ or $C_3H_6O_3$

Answer: **$C_3H_6O_3$**

Marks 4

- If 50 mL of a 0.10 M solution of $AgNO_3$ is mixed with 50 mL of a 0.40 M solution of Na_2CO_3, what mass of Ag_2CO_3 will precipitate from the reaction?

The ionic equation for the precipitation reaction is:

$$2Ag^{+}(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3(s)$$

Thus, two moles of $Ag^{+}(aq)$ are required for every one mole of $CO_3^{2-}(aq)$.

The number of moles of $Ag^{+}(aq)$ and $CO_3^{2-}(aq)$ are given by:

$$n(Ag^{+}(aq)) = \text{concentration} \times \text{volume} = 0.10 \times \frac{50}{1000} = 0.0050 \text{ mol}$$

$$n(CO_3^{2-}(aq)) = 0.40 \times \frac{50}{1000} = 0.020 \text{ mol}$$

There is insufficient $Ag^{+}(aq)$ to react with all of the $CO_3^{2-}(aq)$ and so it is $Ag^{+}(aq)$ is the limiting reagent. From the chemical equation, 1 mole of $Ag_2CO_3(s)$ is produced from every two moles of $Ag^{+}(aq)$ ions. The amount of $Ag_2CO_3(s)$ produced is therefore:

$$n(Ag_2CO_3(s)) = \tfrac{1}{2} \times n(Ag^{+}(aq)) = \tfrac{1}{2} \times 0.0050 = 0.0025 \text{ mol}$$

The formula mass of Ag_2CO_3 is (2 × 107.87 (Ag)) + 12.01 (C) + (3 × 16.00 (O)) = 275.75. This number of moles thus corresponds to a mass of:

$$\text{mass of } Ag_2CO_3 = \text{number of moles} \times \text{formula mass} = 0.0025 \times 275.75 = 0.69 \text{ g}$$

Answer: **0.69 g**

What is the final concentration of CO_3^{2-} ions in the solution after the above reaction?

From the chemical equation, one mole of $Ag_2CO_3(s)$ is produced from every mole of CO_3^{2-} which reacts. Therefore 0.0025 mol of CO_3^{2-} reacts. This leaves:

$$\text{number of moles of unreacted } CO_3^{2-} = 0.020 - 0.0025 = 0.018 \text{ mol}$$

The total volume of the solution after mixing is (50 + 50) = 100 mL. The final concentration is therefore:

$$\text{concentration} = \frac{\text{number of moles}}{\text{volume}} = \frac{0.018}{100/1000} = 0.18 \text{ M}$$

Answer: **0.18 M**

ANSWER CONTINUES ON THE NEXT PAGE

• Give balanced ionic equations for the reactions that occur in each of the following cases.	**3**
Sodium metal is added to excess water. **$2Na(s) + 2H_2O(l) \rightarrow 2Na^+(aq) + 2OH^-(aq) + H_2(g)$**	
Solutions of cobalt(II) nitrate and sodium phosphate are mixed. **$3Co^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Co_3(PO_4)_2(s)$**	
Solid calcium carbonate is dissolved in dilute nitric acid. **$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$**	

Marks

- A 60.0 g piece of Ag metal is heated to 90.0 °C and dropped into 120.0 g of water at 25.0 °C in a well insulated container. The final temperature of the Ag-H_2O mixture is 26.7 °C. Calculate the specific heat of silver.
  Data: The specific heat of water is 4.18 J g^{-1} K^{-1}. 3

The temperature of the Ag metal decreases from 90.0 °C to 26.7 °C. The heat change is given by:

$$\textbf{heat change of Ag} = q_{Ag} = c_{Ag} \times m_{Ag} \times \Delta T = c_{Ag} \times (60.0) \times (26.7 - 90.0)$$

The temperature of the water changes from 25.0 °C to 26.7 °C. The heat change is given by:

$$\textbf{heat change of water} = q_{water} = c_{water} \times m_{water} \times \Delta T = (4.18) \times (120.0) \times (26.7 - 25.0)$$

When the Ag metal is dropped into the water, the heat lost by the silver is gained by the water until both reach the same temperature: $-q_{Ag} = q_{water}$:

$$-c_{Ag} \times (60.0) \times (26.7 - 90.0) = (4.18) \times (120.0) \times (26.7 - 25.0)$$

$$c_{Ag} = \frac{(4.18) \times (120.0) \times (26.7 - 25.0)}{(60.0) \times (90.0 - 26.7)} = 0.22 \text{ J g}^{-1} \text{ K}^{-1}$$

Answer: **0.22 J g^{-1} K^{-1}**

- Determine K_c for the reaction $\frac{1}{2}O_2(g) + Na_2O(s) \rightleftharpoons Na_2O_2(s)$ at 25 °C. 2

  Data: $Na_2O(s) \rightleftharpoons 2Na(s) + \frac{1}{2}O_2(g)$ $K_c = 2 \times 10^{-25}$ at 25 °C.

  $Na_2O_2(s) \rightleftharpoons 2Na(s) + O_2(g)$ $K_c = 5 \times 10^{-29}$ at 25 °C.

The reaction involves the formation of $Na_2O_2(s)$ from $Na_2O(s)$ and thus involves the first reaction and the reverse of the second reaction. The reactions can be combined:

$$Na_2O(s) \rightleftharpoons 2Na(s) + \tfrac{1}{2} O_2(s) \qquad K_c = 2 \times 10^{-25}$$

$$2Na(s) + O_2(g) \rightleftharpoons Na_2O_2(s) \qquad K_c = \frac{1}{(5 \times 10^{-29})} = 2 \times 10^{-30}$$

$$Na_2O(s) + \tfrac{1}{2} O_2(g) \rightleftharpoons Na_2O_2(g) \qquad K_c = (2 \times 10^{-25}) \times \frac{1}{(5 \times 10^{-29})} = 4 \times 10^{3}$$

Answer: **4 × 10^3**

Marks

- Consider a cell composed of the following half-reactions. **2**

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

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

What is the balanced equation for the spontaneous reaction?

The E^{o} values for the Ag^{+}/Ag and Cr/Cr^{3+} cells are +0.80 and -0.74 V respectively. The least positive (Cr/Cr^{3+}) is reversed and is the oxidation half cell:

$$\mathbf{3Ag^{+}(aq) + Cr(s) \rightarrow 3Ag(s) + Cr^{3+}(aq)}$$

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

$$\mathbf{E^{o}_{cell} = E^{o}_{red} + E^{o}_{ox} = (+0.80) + (0.74) = +1.54\,V}$$

The potential for the Cr/Cr^{3+} cell is reversed as it is the oxidation half cell.

Answer: **+1.54 V**

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

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

Data: $\Delta H^{o}_{f} = +51.9$ kJ mol^{-1} for $Cu^{+}(aq)$

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

Using $\Delta_{rxn}H^{o} = \sum m\Delta_{f}H^{o}(\text{products}) - \sum n\Delta_{f}H^{o}(\text{reac tan ts})$, the heat of the reaction as written is:

$$\mathbf{\Delta_{rxn}H^{o} = [2\Delta_{f}H^{o}(Cu(s)) + \Delta_{f}H^{o}(Zn^{2+}(aq))] - [\Delta_{f}H^{o}(Zn(s)) + 2\Delta_{f}H^{o}(Cu^{+}(a}$$

$$\mathbf{= [(2\times 0) + (-152.4)] - [(0) + (2\times +51.9)] = -256.2\,kJ\,mol^{-1}}$$

$\Delta_{f}H^{o}(Cu(s)) = \Delta_{f}H^{o}(Zn(s)) = 0$ for elements already in their standard states.

Answer: **-256.2 kJ mol^{-1}**

Marks

- An electrolytic cell contains a solution of MCl_3. A total charge of 3600 C is passed through the cell, depositing 0.65 g of the metal, M, at the cathode. What is the identity of the metal, M? **4**

The total number of electrons passed is given by:

$$\textbf{number of moles of electrons} = \frac{It}{F} = \frac{Q}{F} = \frac{3600}{96485} = 0.037\ \text{mol.}$$

As the solution contains MCl_3, these electrons are reducing $M^{3+}(aq)$ ions, each of which requires $3e^-$. Hence, the number of moles of metal, M, formed at the cathode is:

number of moles of M = ⅓ × 0.037 = 0.012

This amount of M has a mass of 0.65 g. The atomic mass of M is therefore:

$$\textbf{atomic mass} = \frac{\text{mass}}{\text{number of moles}} = \frac{0.65\ \text{g}}{0.012\ \text{mol}} = 52\ \text{g mol}^{-1}$$

This atomic mass corresponds to that of chromium (Cr).

Answer: **Chromium (Cr)**

- A metal-metal hydride battery has the following shorthand notation: **3**

$$MH(s), M(s)\ |\ OH^-(aq)\ ||\ OH^-(aq)\ |\ NiO(OH)(s), Ni(OH)_2(s)$$

Which component of the battery is the cathode?

$NiO(OH)(s)/Ni(OH)_2(s)$ (the right hand side of the cell notation)

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

$NiO(OH)(s) + H_2O(l) + e^- \rightarrow Ni(OH)_2(s) + OH^-(aq)$

Why is it important that all redox active species are solids in this reaction?

As all redox active species are solids, their concentrations do not change during the reaction. They therefore do not appear in the equilibrium expression or in the Nernst equation.

The battery therefore maintains a constant voltage.

	Marks
• $K_p = 7.0$ for the reaction $Br_2(g) + Cl_2(g) \rightleftharpoons 2BrCl(g)$ at 400 K. Suppose a 1.0 L flask is filled with 0.30 atm $Br_2(g)$ and 0.30 atm $Cl_2(g)$ at 400 K. Find the pressures of all three gases at equilibrium.	4

The equilibrium constant in terms of partial pressures, K_p, is given by:

$$K_p = \frac{(P_{BrCl})^2}{(P_{Br_2})(P_{Cl_2})}$$

The reaction table is:

pressure	$Br_2(g)$	$Cl_2(g)$	$\rightleftharpoons$	$2BrCl(g)$
start	0.30	0.30		0
change	-x	-x		+2x
equilibrium	0.30-x	0.30-x		2x

The equilibrium constant in terms of partial pressures, K_p, is given by:

$$K_p = \frac{(P_{BrCl})^2}{(P_{Br_2})(P_{Cl_2})} = \frac{(2x)^2}{(0.30-x)(0.30-x)} = \frac{(2x)^2}{(0.30-x)^2} = 7.0$$

As K_p is not small, the assumption that 0.30 –x ~ x *cannot* be made. Hence,

$$\frac{(2x)}{(0.30-x)} = \sqrt{7.0}$$

$$2x = (0.30-x)\times\sqrt{7.0} = (0.30\times\sqrt{7}) - x\sqrt{7}$$

$$x = \frac{(0.30\times\sqrt{7})}{(2+\sqrt{7})} = 0.17$$

The partial pressures at equilibrium are therefore:

$$P_{Br_2} = P_{Cl_2} = 0.30 - x = 0.13 \text{ atm}, \; P_{BrCl} = 2x = 0.34 \text{ atm}$$

$p(Br_2)$: **0.13 atm**	$p(Cl_2)$: **0.13 atm**	$p(BrCl)$: **0.34 atm**

Marks

- The *Voyager I* spacecraft determined that the atmospheric pressure at the surface of Saturn's moon, Titan, is 1.6 times that of earth and that the atmosphere contains 6.0 mol % methane, CH_4. What is the partial pressure of methane on Titan in mmHg? **2**

If the atmospheric pressure on Titan is 1.6 times that on earth, P_{total} = 1.6 atm.

If the atmosphere contains 6.0 mol % methane, then:

$$P_{CH_4} = 0.06 \times 1.6 = 0.096 \text{ atm} = 0.096 \times 760 = 73 \text{ mmHg}$$

Answer: **73 mmHg**

- Many gases are available for use in compressed gas cylinders, in which they are stored at high pressures. Calculate the mass of O_2 that can be stored at 20. °C and 170. atm pressure in a cylinder with a volume of 60.0 L. **3**

Using the ideal gas law, PV = nRT, the number of moles that can be stored is:

$$n = \frac{PV}{RT} = \frac{(170.)\times(60.0)}{(0.08206)\times(20.+273)} = 424 \text{ mol}$$

As the molar mass of O_2 is (2 × 16.00) = 32.00, this corresponds to a mass of:

mass = number of moles × molar mass = 424 × 32.00 = 13600 g = 13.6 kg

Answer: **13,600 g or 13.6 kg**

What volume would this mass of oxygen occupy at 1.00 atm pressure and 20 °C?

Using the ideal gas law, the volume of 424 mol at 1.00 atm and 20 °C is:

$$V = \frac{nRT}{P} = \frac{(424)\times(0.08206)\times(20+273)}{(1.00)} = 10.2\times10^4 \text{ L}$$

Equivalently, $P_1V_1 = P_2V_2$ gives:

$$V_2 = \frac{P_1V_1}{P_2} = \frac{(170)\times(60.0)}{(1.00)} = 1.02\times10^4 \text{ L}$$

Answer: **1.02 × 10^4 L**