

Topics in the June 2010 Exam Paper for CHEM1001

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- [Lewis Model of Bonding](#)
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- Complete the following table.

Name	Formula
ammonia	NH ₃
phosphorus trichloride	PCl₃
potassium hydrogencarbonate	KHCO ₃
calcium phosphate	Ca₃(PO₄)₂

Marks
2

- Calculate the number of aluminium atoms in a block of pure aluminium that measures 2.0 cm × 2.0 cm × 3.0 cm. The density of aluminium is 2.7 g cm⁻³.

2

The volume of the block is:

$$V = \text{length} \times \text{width} \times \text{height} = (2.0 \times 2.0 \times 3.0) \text{ cm}^3 = 12 \text{ cm}^3$$

The mass can then be calculated from the density:

$$\text{density} = \text{mass} / \text{volume} \text{ or } \text{mass} = \text{density} \times \text{volume}$$

$$\text{mass} = (2.7 \text{ g cm}^{-3}) \times (12.0 \text{ cm}^3) = 32.4 \text{ g}$$

1 mol of Al has a mass equal to its atomic mass, 26.98 g mol⁻¹ and contains 6.022 × 10²³ mol⁻¹. Hence, the number of atoms in 32.4 g is:

$$\begin{aligned} \text{number of atoms} &= \text{number of moles} \times \text{Avogadro's number} \\ &= (32.4 \text{ g} / 26.98 \text{ g mol}^{-1}) \times (6.022 \times 10^{23} \text{ mol}^{-1}) = 7.2 \times 10^{23} \end{aligned}$$

Answer: **7.2 × 10²³**

- Complete the following table.

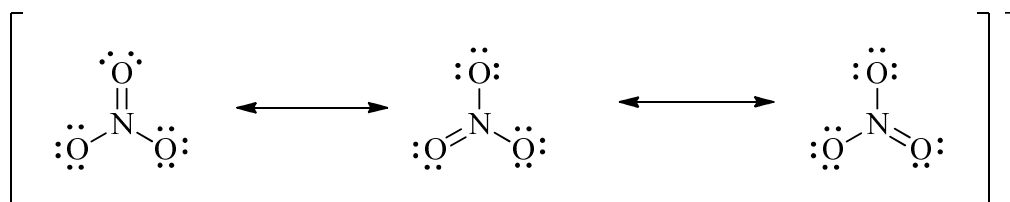
Marks
6

Molecule	NH ₃	SF ₄	CO ₂
Number of bonding electron pairs	3	4	4
Number of non-bonding electron pairs on central atom	1	1	0
Lewis structure	$\begin{array}{c} \text{H} - \overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{N}}} - \text{H} \\ \\ \text{H} \end{array}$	$\begin{array}{c} \text{:F:} \\ \text{:F:} \diagdown \quad \quad \diagup \text{:F:} \\ \text{:F:} \quad \text{S} \quad \text{:F:} \\ \text{:F:} \quad \\ \text{:F:} \end{array}$	$\text{:}\ddot{\text{O}}=\text{C}=\ddot{\text{O}}\text{:}$
Molecular shape	trigonal pyramidal	“see saw”	linear

- What is resonance? Illustrate your answer by using the nitrate ion, NO₃⁻, as an example.

3

When two or more Lewis structures can be drawn for a molecule, the true structure is none of the structures that is drawn, but a type of average made up of all the resonance contributors. Some structures may contribute more than others.



In NO₃⁻, the ion does not contain 1 double and 2 single bonds, but is an average of the three structures shown. All of the N-O bonds are exactly the same length and the energy of the true structure is lower than the theoretical energy for any one of the given structures. This energy difference is known as resonance stabilisation energy.

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- The element titanium is used as a structural material for bone in joint replacement surgery. Discuss the properties of titanium that make it suitable for this application.

Marks
3

Titanium is inert. It will not corrode nor dissolve and poison the patient.
Titanium is a light-weight metal with a similar density to bone.
Titanium is strong and will not break, so is suitable for supporting the patient's weight.

- Gaseous lithium atoms absorb light with a wavelength of 323 nm. The resulting excited lithium atoms lose some energy through collisions with other atoms. One of the emission lines has an energy of 2.44×10^{-19} J.

4

Calculate the energy of the light used for the excitation.

The energy of electromagnetic radiation can be calculated from its wavelength using Planck's equation:

$$E = hc / \lambda$$

As the absorbed light has $\lambda = 323$ nm:

$$E = (6.626 \times 10^{-34} \text{ J s})(2.998 \times 10^8 \text{ m s}^{-1}) / (323 \times 10^{-9} \text{ m}) = 6.15 \times 10^{-19} \text{ J}$$

Answer: 6.15×10^{-19} J

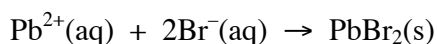
Calculate the wavelength of the light emitted.

The emitted light has $E = 2.44 \times 10^{-19}$ J and the same relationship can be used to calculate the corresponding wavelength:

$$\begin{aligned} \lambda &= hc / E \\ &= (6.626 \times 10^{-34} \text{ J s})(2.998 \times 10^8 \text{ m s}^{-1}) / (2.44 \times 10^{-19} \text{ J}) \\ &= 8.14 \times 10^{-7} \text{ m} = 814 \text{ nm} \end{aligned}$$

Answer: 814 nm

- Lead ions react with bromide ions according to the following equation.



If 0.040 M lead(II) nitrate solution (100.0 mL) is added to 0.020 M potassium bromide solution (300.0 mL), what amount (in mol) of lead(II) bromide precipitates?

Marks
4

The number of moles of Pb^{2+} ions in 100.0 mL of a 0.040 M solution of $\text{Pb}(\text{NO}_3)_2$ is:

$$\begin{aligned} \text{number of moles} &= \text{concentration} \times \text{volume} \\ &= (0.040 \text{ mol L}^{-1}) \times (0.1000 \text{ L}) = 0.0040 \text{ mol} \end{aligned}$$

The number of moles of Br^{-} ions in 300.0 mL of a 0.020 M solution of KBr is:

$$\text{number of moles} = (0.020 \text{ mol L}^{-1}) \times (0.3000 \text{ L}) = 0.0060 \text{ mol}$$

The precipitation reaction requires 2 Br^{-} ions for every Pb^{2+} ion. As there is *less* than twice as much Br^{-} than Pb^{2+} , it is the Br^{-} that is the limiting reagent.

From the chemical equation, 2 mol of Br^{-} leads to 1 mol of $\text{PbBr}_2(\text{s})$ and so:

$$\text{number of moles of } \text{PbBr}_2(\text{s}) \text{ formed} = \frac{1}{2} \times 0.0060 \text{ mol} = 0.0030 \text{ mol}$$

Answer: **0.0030 mol**

What is the final concentration of $\text{NO}_3^{-}(\text{aq})$ ions remaining in solution after the reaction?

NO_3^{-} is not involved in the reaction: it is a spectator ion. The amount present after the reaction is the same as at the beginning. When 1 mol of $\text{Pb}(\text{NO}_3)_2$ dissolves, it forms 1 mol of $\text{Pb}^{2+}(\text{aq})$ and 2 mol of $\text{NO}_3^{-}(\text{aq})$. Hence, the number of moles of $\text{NO}_3^{-}(\text{aq})$ present is:

$$\text{number of moles} = 2 \times (0.040 \text{ mol L}^{-1}) \times (0.1000 \text{ L}) = 0.0080 \text{ mol}$$

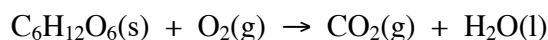
After the solutions are mixed, the total volume is (100.0 + 300.0) mL = 400.0 mL.

This amount is now present in this volume and so has a concentration:

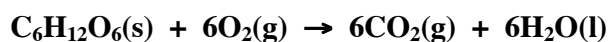
$$\text{concentration} = \text{number of moles} / \text{volume} = 0.0080 \text{ mol} / 0.4000 \text{ L} = 0.020 \text{ M}$$

Answer: **0.020 M**

- Respiration involves the oxidation of glucose to produce carbon dioxide, water and energy:



Balance this equation.



What volume of $\text{CO}_2(\text{g})$ is produced from the oxidation of 10.0 g of glucose under body conditions (37°C , $1.00 \times 10^5 \text{ Pa}$)?

The molar mass of glucose is:

$$\text{molar mass} = (6 \times 12.01 + 12 \times 1.008 + 6 \times 16.00) \text{ g mol}^{-1} = 180.156 \text{ g mol}^{-1}$$

Therefore, 10.0 g corresponds to:

$$\begin{aligned} \text{number of moles of glucose} &= \text{mass} / \text{molar mass} \\ &= 10.0 \text{ g} / 180.156 \text{ g mol}^{-1} = 0.0555 \text{ mol} \end{aligned}$$

From the balanced chemical equation, 1 mol of glucose leads to 6 mol of $\text{CO}_2(\text{g})$ so:

$$\text{number of moles of } \text{CO}_2(\text{g}) \text{ generated} = 6 \times 0.0555 \text{ mol} = 0.333 \text{ mol}$$

From the ideal gas law, $PV = nRT$ with $R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$ when the pressure is used in the Pascals (the S.I. unit). The volume that this occupies is:

$$\begin{aligned} V &= nRT / P \\ &= (0.333 \text{ mol}) \times (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times ((37 + 273) \text{ K}) / (1.00 \times 10^5 \text{ Pa}) \\ &= 8.58 \times 10^{-3} \text{ m}^3 = 8.58 \text{ L} \end{aligned}$$

Note that as S.I. units were used, the volume is returned in m^3 requiring the conversion to litres.

Alternatively, the pressure can be converted from Pa to atm:

$$P = 1.00 \times 10^5 \text{ Pa} = (1.00 \times 10^5 / 101.3 \times 10^3) = 0.987 \text{ atm}$$

From the ideal gas law, $PV = nRT$ with $R = 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}$ when the pressure is used in the atmospheres. The volume that this occupies is:

$$\begin{aligned} V &= nRT / P \\ &= (0.333 \text{ mol}) \times (0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}) \times ((37 + 273) \text{ K}) / (0.987 \text{ atm}) \\ &= 8.58 \text{ L} \end{aligned}$$

Answer: **8.58 L**

Marks
4

- An unknown liquid contains H: 5.90 % and O: 94.1 % by mass and has a molar mass of 33.9 g mol^{-1} . What is its molecular formula?

Marks
2

The liquid contains 5.90% H and so 94.1% O.

	H	O
percentage	5.90	94.1
divide by atomic mass	$\frac{5.90}{1.008} = 5.85$	$\frac{94.1}{16.00} = 5.88$
divide by smallest value	1	1

The ratio of H : O is 1 : 1 and so the empirical formula is HO.

The molecular formula is $(\text{HO})_n$. The molar mass is:

$$\text{molar mass} = n \times (1.008 + 16.00) \text{ g mol}^{-1} = 17.008n \text{ g mol}^{-1}$$

As the molar mass is 33.9 g mol^{-1} , $n = 2$ and the molecular formula is $(\text{HO})_2$ or H_2O_2 . It is hydrogen peroxide.

Answer: H_2O_2

- A 2.4 g sample of zinc was dropped into 0.250 L of 5.0 M HCl in a 5.00 L container at 25°C with an initial pressure of 1.0 atm and then the vessel sealed. Calculate the final pressure inside the container.
Hint: The volume occupied by the HCl is significant.

6

The reaction occurring is:



The number of moles of zinc added is:

$$\text{number of moles of Zn(s)} = \text{mass} / \text{molar mass} = 2.4 \text{ g} / 65.39 \text{ g mol}^{-1} = 0.037 \text{ mol}$$

The number of moles of $\text{H}^+(\text{aq})$ initially present is:

$$\begin{aligned} \text{number of moles of H}^+(\text{aq}) &= \text{concentration} \times \text{volume} \\ &= 5.0 \text{ mol L}^{-1} \times 0.250 \text{ L} = 1.25 \text{ L} \end{aligned}$$

As 2 mol of $\text{H}^+(\text{aq})$ is required for every 1 mol of Zn(s) , the $\text{H}^+(\text{aq})$ is in excess and zinc is the limiting reagent. From the chemical equation, 1 mol of Zn(s) produces 1 mol of $\text{H}_2(\text{g})$ and so:

$$\text{number of moles of H}_2(\text{g}) = 0.037 \text{ mol}$$

ANSWER CONTINUES ON THE NEXT PAGE

This gas is contained within 4.75 L of the container as the HCl solution occupies 0.250 L. Using the ideal gas law, $PV = nRT$ with $R = 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}$, this amount will lead to a pressure:

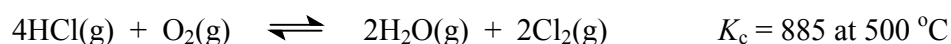
$$\begin{aligned} P &= nRT / V \\ &= (0.037 \text{ mol}) \times (0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}) \times ((25 + 273) \text{ K}) / (4.75 \text{ L}) \\ &= 0.19 \text{ atm} \end{aligned}$$

As the initial pressure was 1.0 atm, the total pressure in the container has increased to:

$$\text{total pressure} = (1.0 + 0.19) \text{ atm} = 1.2 \text{ atm}$$

Answer: **1.2 atm**

- Consider the following equilibrium reaction.



If 0.030 mol HCl, 0.020 mol O₂, 0.090 mol H₂O and 0.085 mol Cl₂ are mixed in a 1.0 L container at 500 °C, in what direction will the reaction proceed?

Marks
3

As the gases are present in a 1.0 L container, their concentrations are:

concentration = number of moles / volume

$$[\text{HCl}] = 0.030 \text{ mol} / 1.0 \text{ L} = 0.030 \text{ M}$$

$$[\text{O}_2] = 0.020 \text{ mol} / 1.0 \text{ L} = 0.020 \text{ M}$$

$$[\text{H}_2\text{O}] = 0.090 \text{ mol} / 1.0 \text{ L} = 0.090 \text{ M}$$

$$[\text{Cl}_2] = 0.085 \text{ mol} / 1.0 \text{ L} = 0.085 \text{ M}$$

The reaction quotient, Q , is:

$$Q = \frac{[\text{H}_2\text{O}(\text{g})]^2[\text{Cl}_2(\text{g})]^2}{[\text{HCl}(\text{g})]^4[\text{O}_2(\text{g})]} = \frac{(0.090)^2[0.085]^2}{[0.030]^4[0.020]} = 3600$$

As $Q > K_c$, the reaction will proceed towards the reactants.

Answer: **towards reactants**

What is the value of K_p for the reaction at 500 °C?

The reaction involves 5 mol of gaseous reactants going to 4 mol of gaseous products. The number of moles of gas decreases in the reaction with $\Delta n = -1$.

K_p and K_c are related by:

$$K_p = K_c(RT)^{\Delta n}$$

Hence,

$$K_p = (885) \times [0.08206 \times (500 + 273)]^{-1} = 14$$

Answer: **14**

- How much energy is needed to convert 15 g of ice at 0.0 °C to water at 60.0 °C? The molar heat of fusion of water is 6.009 kJ mol⁻¹ and the specific heat capacity of water is 4.18 J g⁻¹ K⁻¹.

Marks
3

The molar mass of H₂O is $(2 \times 1.008 + 16.00) \text{ g mol}^{-1} = 18.018 \text{ g mol}^{-1}$. Hence, 15 g contains:

$$\text{number of moles} = \text{mass} / \text{molar mass} = 15 \text{ g} / 18.018 \text{ g mol}^{-1} = 0.83 \text{ mol.}$$

Energy is required to (i) melt the ice and (ii) heat it;

- (i) 6.009 kJ is required to melt 1 mol. Hence, to melt 0.83 mol would require:

$$q_{\text{melt}} = (0.83 \text{ mol}) \times (6.009 \text{ kJ mol}^{-1}) = 5.0 \text{ kJ}$$

- (ii) The heat required to heat water from 0.0 °C to 60.0 °C is given by:

$$q_{\text{heat}} = mC\Delta T = (15 \text{ g}) \times (4.18 \text{ J g}^{-1} \text{ K}^{-1}) \times ((60.0 - 0.0) \text{ K}) = 3800 \text{ J}$$

Overall;

$$q_{\text{total}} = q_{\text{melt}} + q_{\text{heat}} = (5.0 \times 10^3 \text{ J}) + (3800 \text{ J}) = 8800 \text{ J} = 8.8 \text{ J}$$

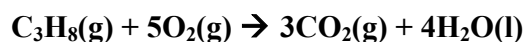
Answer: **8.8 J**

- Calculate the standard enthalpy change for the combustion of 1.00 mol of propane gas, C₃H₈(g), to CO₂(g) and H₂O(l).

3

compound	C ₃ H ₈ (g)	CO ₂ (g)	H ₂ O(l)
$\Delta_f H^\circ / \text{kJ mol}^{-1}$	-105	-394	-286

The chemical equation for the combustion reaction is:



Using $\Delta_{\text{rxn}} H^\circ = \sum m \Delta_f H^\circ(\text{products}) - \sum n \Delta_f H^\circ(\text{reactants})$,

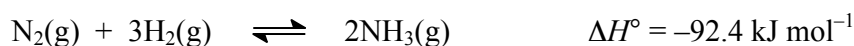
$$\Delta_{\text{rxn}} H^\circ = [3 \times \Delta_f H^\circ(\text{CO}_2(\text{g})) + 4 \times \Delta_f H^\circ(\text{H}_2\text{O}(\text{l}))] - [\Delta_f H^\circ(\text{C}_3\text{H}_8(\text{g}))]$$

$$= [(3 \times -394) + (4 \times -286)] - [-105] \text{ kJ mol}^{-1}$$

$$= -2220 \text{ kJ mol}^{-1}$$

Answer: **-2220 kJ mol⁻¹**

- Ammonia can be produced according to the following equation.



$K_c = 1.5 \times 10^{-5}$ at 500 K. What is the concentration of ammonia at equilibrium if 0.5 mol of $\text{N}_2(\text{g})$ and 1.5 mol of $\text{H}_2(\text{g})$ are placed in an empty 2.0 L flask and allowed to come to equilibrium at 500 K?

Marks
6

The initial concentrations of $\text{N}_2(\text{g})$ and $\text{H}_2(\text{g})$ are:

$$[\text{N}_2(\text{g})] = \text{number of moles} / \text{volume} = 0.5 \text{ mol} / 2.0 \text{ L} = 0.25 \text{ M}$$

$$[\text{H}_2(\text{g})] = 1.5 \text{ mol} / 2.0 \text{ L} = 0.75 \text{ M}$$

A reaction table can be used to work out the equilibrium concentrations:

	$\text{N}_2(\text{g})$	$3\text{H}_2(\text{g})$		$2\text{NH}_3(\text{g})$
initial	0.25	0.75	\rightleftharpoons	0
change	-x	-3x		+2x
equilibrium	0.25 - x	0.75 - 3x		2x

The equilibrium constant, K_c , is;

$$K_c = \frac{[\text{NH}_3(\text{g})]^2}{[\text{N}_2(\text{g})][\text{H}_2(\text{g})]^3} = \frac{(2x)^2}{(0.25-x)(0.75-3x)^3} = 1.5 \times 10^{-5}$$

The equilibrium constant is *very* small so x is tiny. As a result, $0.25 - x \approx 0.25$ and $0.75 - 3x \approx 0.75$. With this approximation:

$$K_c = \frac{(2x)^2}{(0.25)(0.75)^3} = 1.5 \times 10^{-5}$$

$$4x^2 = 1.5 \times 10^{-5} \times 0.105 \quad \text{and so } x = 0.00063 \text{ mol.}$$

Hence, $[\text{NH}_3(\text{g})] = 2x = 1.3 \times 10^{-3} \text{ M} = 1 \times 10^{-3} \text{ M}$ (1 sf)

Answer: $1.3 \times 10^{-3} \text{ M} = 1 \times 10^{-3} \text{ M}$ (1 sf)

An additional 0.5 mol of nitrogen is added to the flask described above and equilibrium re-established. Will the equilibrium constant have increased, decreased or remained the same? Justify your answer.

It remains the same. As the name implies, it is constant and is unaffected by increases in amounts of reactants or products.

What now is the equilibrium concentration of ammonia?

Before addition of the additional nitrogen, the concentrations are (from above):

$$[\text{N}_2(\text{g})] = 0.25 \text{ M}, [\text{H}_2(\text{g})] = 0.75 \text{ M and } [\text{NH}_3(\text{g})] = 1.3 \times 10^{-3} \text{ M}$$

The amount of $\text{N}_2(\text{g})$ is given by;

$$\begin{aligned} \text{number of moles of } \text{N}_2(\text{g}) &= \text{concentration} \times \text{volume} \\ &= 0.25 \text{ mol L}^{-1} \times 2.0 \text{ L} = 0.50 \text{ mol} \end{aligned}$$

After addition of 0.5 mol, the amount is therefore $(0.50 + 0.5) \text{ mol} = 1.0 \text{ mol}$. The concentration is therefore:

$$[\text{N}_2(\text{g})] = \text{number of moles} / \text{volume} = 1.0 \text{ mol} / 2.0 \text{ L} = 0.50 \text{ M}$$

Another reaction table can be used to work out the new equilibrium concentrations:

	$\text{N}_2(\text{g})$	$3\text{H}_2(\text{g})$	\rightleftharpoons	$2\text{NH}_3(\text{g})$
initial	0.50	0.75		1.3×10^{-3}
change	-y	-3y		+2y
equilibrium	$0.50 - y$	$0.75 - 3y$		$1.3 \times 10^{-3} + 2y$

$$K_c = \frac{[\text{NH}_3(\text{g})]^2}{[\text{N}_2(\text{g})][\text{H}_2(\text{g})]^3} = \frac{(1.3 \times 10^{-3} + 2y)^2}{(0.50 - y)(0.75 - 3y)^3} = 1.5 \times 10^{-5}$$

The equilibrium constant is *very* small so y is also tiny. As a result, $0.50 - y \approx 0.50$ and $0.75 - 3y \approx 0.75$. With this approximation:

$$K_c = \frac{(1.3 \times 10^{-3} + 2y)^2}{(0.50)(0.75)^3} = 1.5 \times 10^{-5}$$

$$(1.3 \times 10^{-3} + 2y)^2 = 1.5 \times 10^{-5} \times 0.211 \quad \text{and so } y = 0.00024 \text{ mol.}$$

$$[\text{NH}_3(\text{g})] = 1.3 \times 10^{-3} + 2y = 1.8 \times 10^{-3} \text{ M}$$

Note that $[\text{NH}_3(\text{g})]$ has increased, which is consistent with Le Chatelier's principle as more reactant has been added.

Answer: $1.8 \times 10^{-3} \text{ M}$

ANSWER CONTINUES ON THE NEXT PAGE

What now is the equilibrium concentration of ammonia?

Before addition of the additional nitrogen, the concentrations are (from above):

$$[\text{N}_2(\text{g})] = 0.25 \text{ M}, [\text{H}_2(\text{g})] = 0.75 \text{ M and } [\text{NH}_3(\text{g})] = 1.3 \times 10^{-3} \text{ M}$$

The amount of $\text{N}_2(\text{g})$ is given by;

$$\begin{aligned} \text{number of moles of } \text{N}_2(\text{g}) &= \text{concentration} \times \text{volume} \\ &= 0.25 \text{ mol L}^{-1} \times 2.0 \text{ L} = 0.50 \text{ mol} \end{aligned}$$

After addition of 0.5 mol, the amount is therefore $(0.50 + 0.5) \text{ mol} = 1.0 \text{ mol}$. The concentration is therefore:

$$[\text{N}_2(\text{g})] = \text{number of moles} / \text{volume} = 1.0 \text{ mol} / 2.0 \text{ L} = 0.50 \text{ M}$$

Another reaction table can be used to work out the new equilibrium concentrations:

	$\text{N}_2(\text{g})$	$3\text{H}_2(\text{g})$	\rightleftharpoons	$2\text{NH}_3(\text{g})$
initial	0.50	0.75		1.3×10^{-3}
change	-y	-3y		+2y
equilibrium	$0.50 - y$	$0.75 - 3y$		$1.3 \times 10^{-3} + 2y$

$$K_c = \frac{[\text{NH}_3(\text{g})]^2}{[\text{N}_2(\text{g})][\text{H}_2(\text{g})]^3} = \frac{(1.3 \times 10^{-3} + 2y)^2}{(0.50 - y)(0.75 - 3y)^3} = 1.5 \times 10^{-5}$$

The equilibrium constant is *very* small so y is also tiny. As a result, $0.50 - y \approx 0.50$ and $0.75 - 3y \approx 0.75$. With this approximation:

$$K_c = \frac{(1.3 \times 10^{-3} + 2y)^2}{(0.50)(0.75)^3} = 1.5 \times 10^{-5}$$

$$(1.3 \times 10^{-3} + 2y)^2 = 1.5 \times 10^{-5} \times 0.211 \quad \text{and so } y = 0.00024 \text{ mol.}$$

$$[\text{NH}_3(\text{g})] = 1.3 \times 10^{-3} + 2y = 1.8 \times 10^{-3} \text{ M}$$

Note that $[\text{NH}_3(\text{g})]$ has increased, which is consistent with Le Chatelier's principle as more reactant has been added.

Answer: $1.8 \times 10^{-3} \text{ M}$

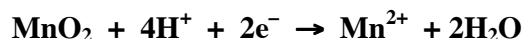
- Write the two half equations and hence balance the equation for the following redox reaction:



Marks
3

Working

The Mn is being reduced since it begins with oxidation number +4 (in MnO_2) and ends with oxidation number +2 (in MnSO_4). The reduction reaction is:

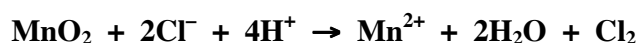


The Cl is being oxidised since it begins with oxidation number -1 (in Cl^-) and ends with oxidation number 0 (in Cl_2). The oxidation reaction is:



These are then summed together to give the overall reaction.

Balanced equation:



Which species is oxidised?

Cl^- is oxidised

- In the electro-refining of Pt, what mass of Pt is deposited from a solution of PtCl_6^{2-} in 1.00 hour, by a current of 1.62 A?

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PtCl_6^{2-} contains Pt(IV) and so 4 moles of electrons are required to reduced 1 mol.

The number of moles of electrons delivered by a current of 1.62 A in 1.00 hour is:

$$\begin{aligned} \text{number of moles of electrons} &= It / F \\ &= (1.62 \text{ C s}^{-1})(1.00 \times 60 \times 60 \text{ s}) / (96485 \text{ C mol}^{-1}) \\ &= 0.0604 \text{ mol} \end{aligned}$$

This will deposit:

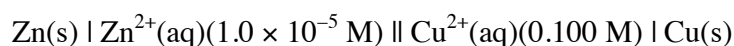
$$\text{number of moles of Pt deposit} = \frac{1}{4} \times 0.0604 \text{ mol} = 0.0151 \text{ mol}$$

As the molar mass of Pt is $195.09 \text{ g mol}^{-1}$, this corresponds to:

$$\begin{aligned} \text{mass of Pt} &= \text{number of moles} \times \text{molar mass} = (0.0151 \text{ mol}) \times (195.09 \text{ g mol}^{-1}) \\ &= 2.95 \text{ g} \end{aligned}$$

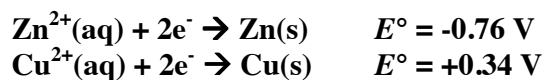
Answer: **2.95 g**

- What is the voltage of the following electrochemical cell at 25 °C?

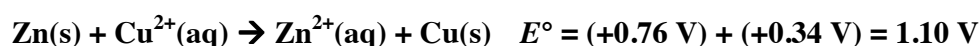


Marks
3

The two half cell reactions and standard reduction potentials are:



As the $\text{Zn}^{2+} / \text{Zn}$ reduction potential is the smallest (i.e. more negative), it is reversed and becomes the oxidation reaction with $E^{\circ} = +0.76 \text{ V}$. The overall reaction and standard cell potential is then:



As non-standard concentrations are used, the Nernst equation must be used to calculate the potential for this 2 electron reaction:

$$\begin{aligned} E &= E^{\circ} - \frac{RT}{nF} \ln Q = E^{\circ} - \frac{RT}{nF} \ln \frac{[\text{Zn}^{2+}(\text{aq})]}{[\text{Cu}^{2+}(\text{aq})]} \\ &= (+1.10 \text{ V}) - \frac{(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(25 + 273 \text{ K})}{(2)(96485 \text{ C mol}^{-1})} \ln \frac{(1.0 \times 10^{-5})}{(0.100)} = +1.22 \text{ V} \end{aligned}$$

Answer: **+1.22 V**

- Rationalise the order of the boiling points of the following liquids in terms of their intermolecular forces.

liquid	F ₂	HCl	HBr	Cl ₂	HF	Br ₂
b.p. (° C)	-188	-85	-67	-34	20	59

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The boiling points in F₂, Cl₂ and Br₂ are determined by the size of the dispersion forces between molecules. The bigger the atoms, the more polarisable their electron clouds and the greater the dispersion forces. Hence boiling points are in order Br₂ > Cl₂ > F₂.

Dispersion forces also operate in HF, HCl and HBr, but here the dipole formed between the halogen atom and the hydrogen also needs to be considered. F is a very small and very electronegative atom. The H-F bond is therefore highly polarised and strong H-bonds form in this liquid. These are much stronger than dispersion forces and so HF has an anomalously high boiling point. Cl and Br are not as electronegative as F: the dispersion forces in HCl and HBr are more significant than the dipole-dipole forces as can be evidenced by the order of boiling points HF > HBr > HCl.

The values given tell us that the total of the dispersion forces in Br₂ is greater than the H-bonds in HF.

- The aluminium-air battery, in which aluminium metal is oxidised to Al^{3+} and O_2 is reduced to OH^- , is being considered as a power source in cars. Briefly compare the relative merits of such a battery with those of a fuel cell for such applications.

Marks
4

Fuel cells use $\text{H}_2(\text{g})$ and $\text{O}_2(\text{g})$ as the reactants.

Disadvantages:

H_2 is highly flammable and a severe explosion hazard should the car be involved in a crash. It's likely cars would need to be heavier to prevent rupture of tanks. There are also handling difficulties as it is a gas.

Advantages: The only product of a fuel cell is water, so they are non-polluting. High efficiency.

Aluminium air battery uses $\text{Al}(\text{s})$ and O_2 as reactants.

Advantages: The fuel (Al) is not explosive and can easily be replaced when it is exhausted. Aluminium is very light metal and 3 electrons lost in its oxidation to Al^{3+} , so a lot of energy generated per g of fuel. Aluminium is very plentiful and relatively cheap. Steady voltage is obtained as $[\text{Al}(\text{s})]$ and $[\text{O}_2(\text{g})]$ do not vary as aluminium is present as a solid and oxygen is obtained from the air.

Disadvantages: The $\text{Al}(\text{OH})_3$ product needs to be recycled. Conversion back to Al metal involves large amounts of electricity and associated CO_2 output.