Topics in the June 2013 Exam Paper for CHEM1001

Click on the links for resources on each topic.

2013-J-2:

- Lewis Model of Bonding
- VSEPR

2013-J-3:

- Lewis Model of Bonding
- Elements and Atoms

2013-J-4:

- Stoichiometry
- Gas Laws
- Molecules and Ions
- The Periodic Table

2013-J-5:

- Elements and Atoms
- Molecules and Ions
- Chemical Equations
- The Periodic Table

2013-J-6:

- Stoichiometry

2013-J-7:

- Lewis Model of Bonding
- Types of Intermolecular Forces

2013-J-8:

- Batteries and Corrosion
- Electrolytic Cells

2013-J-9:

- Gas Laws
- Thermochemistry

2013-J-10:

- Chemical Equilibrium

2013-J-11:

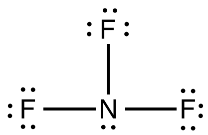
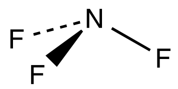
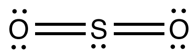
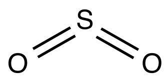
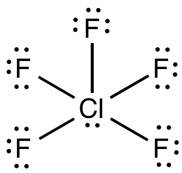
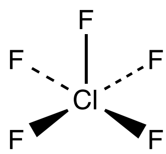
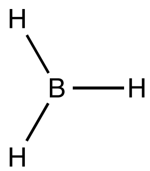
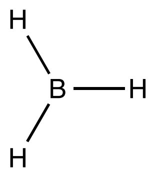
- Types of Intermolecular Forces

2013-J-12:

- Thermochemistry
- First Law of Thermodynamics

Marks
10

- Complete the following table, including resonance structures where appropriate. The central atom is underlined.

Species	Lewis structure	Molecular geometry	Is the species polar?
<u>N</u>F_3		 **Trigonal pyramidal**	**Yes**
<u>S</u>O_2		 **V-shaped**	**Yes**
<u>Cl</u>F_5		 **Square-based pyramidal**	**Yes**
<u>B</u>H_3		 **Trigonal planar**	**No**

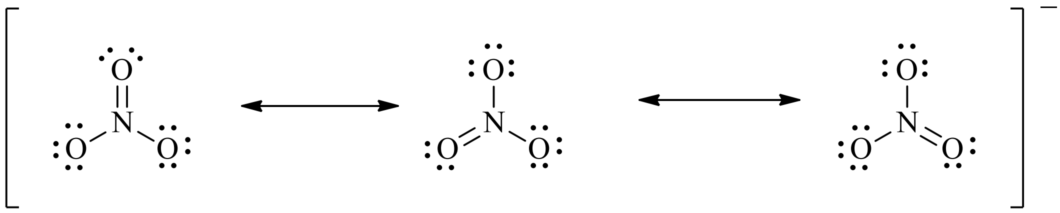
THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY

Marks

• Explain the term 'resonance structures' and give an example. 2

Resonance structures are different valid Lewis structures for a molecule or ion. The true distribution of electrons is an average of all of the different resonance structures.

For example, there are three equivalent resonance structures for the NO_3^- ion:



• Explain why stable compounds of oxygen have 8 electrons in the valence shell, but compounds of sulfur may have 8, 10 or 12 electrons in their valence shell. 2

Oxygen belongs to the 2nd period and there is only space for 8 electrons in its valence shell. Stable compounds of oxygen obey the octet rule.

Sulfur belongs to the 3rd period and there is space for up to 18 electrons in its valence shell. It has 6 valence electrons and can use 2, 4 or 6 of these to make covalent bonds leading to molecules like SF_2, SF_4 and SF_6 respectively. These have 8, 10 and 12 electrons around sulfur respectively.

• In the spaces provided, briefly explain the meaning of the following terms. 3

Valence electrons

The electrons in the outer shell of an atom that can contribute to bonding.

Polar bond

A covalent bond involving two elements of different electronegativity. The different electronegativities lead to an unequal share of the bonding electrons, resulting in a partial positive charge at one end and a partial negative charge at the other end of the bond.

Intensive properties

A physical property of a substance such as density that does not depend on the amount of the substance present.

Marks

- In an experiment, 5.0 g of magnesium was dissolved in excess hydrochloric acid to give magnesium ions and hydrogen gas. Write a balanced equation for the reaction that occurred. **4**

$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

What amount of hydrogen gas (in mol) is produced in the reaction?

The molar mass of Mg is 24.31 g mol^{-1}. 5.0 g therefore corresponds to:

number of moles = mass / molar mass = 5.0 g / 24.31 g mol^{-1} = 0.21 mol

From the chemical equation, each mol of Mg that reacts will give one mol of H_2. Hence,

number of moles of H_2 = 0.21 mol.

Answer: **0.21 mol**

What volume would the hydrogen occupy at 25 °C and 100.0 kPa pressure?

Using the ideal gas law, $PV = nRT$. Hence,

$$\begin{aligned} V &= nRT / P \\ &= (0.21 \text{ mol}) \times (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times ((25 + 273) \text{ K}) / (100.0 \times 10^3 \text{ Pa}) \\ &= 0.0051 \text{ m}^3 \end{aligned}$$

As 1 m^3 = 1000 L, this corresponds to:

V = 1000 × 0.0051 L = 5.1 L

Answer: **5.1 L**

- Silicon and carbon are both in Group 14 and form dioxides. Carbon dioxide is a gas at room temperature while silicon dioxide (sand) is a solid with a high melting point. Describe the bonding in these two materials and explain the differences in properties they show. **3**

CO_2 contains discrete molecules. Carbon makes four bonds by making two C=O double bonds. The C=O double bonds have strong σ and π components. Although these bonds are quite polar, these molecules are linear and do not possess dipole moments. Only very weak dispersion intermolecular forces hold the molecules together and CO_2 is a gas at room temperature.

SiO_2 is a network covalent solid. Each silicon makes four bonds by making four Si-O single bonds. The covalent network leads to a very strongly bonded solid with a very high melting point.

Marks

- Complete the following table by filling in the compound name or formula as required. **2**

Name	Formula
lead(II) chloride	**$PbCl_2$**
dinitrogen trioxide	**N_2O_3**
sodium sulphate	Na_2SO_4
sulfur hexafluoride	SF_6

- In the Periodic Table given, hydrogen is placed at the top of Group 1. List reasons for and against placing hydrogen in this position. **4**

For:

- **It has one valence electron, just like the other members of Group 1**
- **This electron is in a *s*-orbital, just like the other members of Group 1**
- **It has a valency of 1, just like the other members of Group 1**
- **It has low electronegativity, just like the other members of Group 1**

Against:

- **Unlike the other members of Group 1, it is a non-metal**
- **Unlike the other members of Group 1, it exists as diatomic molecules**
- **Unlike the other members of Group 1, it can form both a cation (H^+) and an anion (H^-)**
- **Unlike the other members of Group 1, it is 1 electron short of a noble gas configuration**
- **It forms covalent molecules with non-metals rather than ionic compounds**
- **It is much more electronegative than the other members of Group 1**

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Marks
7

- A 0.060 M solution of aluminium nitrate and a 0.080 M solution of potassium phosphate are prepared by dissolving $Al(NO_3)_3$ and K_3PO_4 in water. Write the ionic equations for these two dissolutions reactions.

Dissolution of $Al(NO_3)_3$	$\mathbf{Al(NO_3)_3(s) \rightarrow Al^{3+}(aq) + 3NO_3^{-}(aq)}$
Dissolution of K_3PO_4	$\mathbf{K_3PO_4(s) \rightarrow 3K^{+}(aq) + PO_4^{3-}(aq)}$

If these solutions are combined, aluminium phosphate precipitates. Write the ionic equation for the precipitation reaction.

$\mathbf{Al^{3+}(aq) + PO_4^{3-}(aq) \rightarrow AlPO_4(s)}$

100.0 mL of the aluminium nitrate solution is added to 50.0 mL of the potassium phosphate solution. What amount (in mol) of aluminium phosphate precipitates?

100.0 mL of a 0.060 M solution of $Al(NO_3)_3$ contains:

number of moles of Al^{3+} = concentration × volume = $c \times V$
= 0.060 mol L^{-1} × 0.1000 L = 0.0060 mol

50.0 mL of a 0.080 M solution of K_3PO_4 contains:

number of moles of PO_4^{3-} = concentration × volume = $c \times V$
= 0.080 mol L^{-1} × 0.0500 L = 0.0040 mol

As the ionic equation has a 1 : 1 ratio of Al^{3+} : PO_4^{3-} reacting, PO_4^{3-} is the limiting reagent. The ionic equation shows that 1 mol of $AlPO_4$ is made from 1 mol of PO_4^{3-} so 0.0040 mol will produce 0.0040 mol.

Answer: **0.0040 mol**

What is the final concentration of aluminium ions remaining in solution after the precipitation?

Formation of 0.0040 mol of $AlPO_4$ requires 0.0040 mol of Al^{3+}. Therefore, the amount remaining is:

number of moles of Al^{3+} remaining = (0.0060 – 0.0040) mol = 0.0020 mol

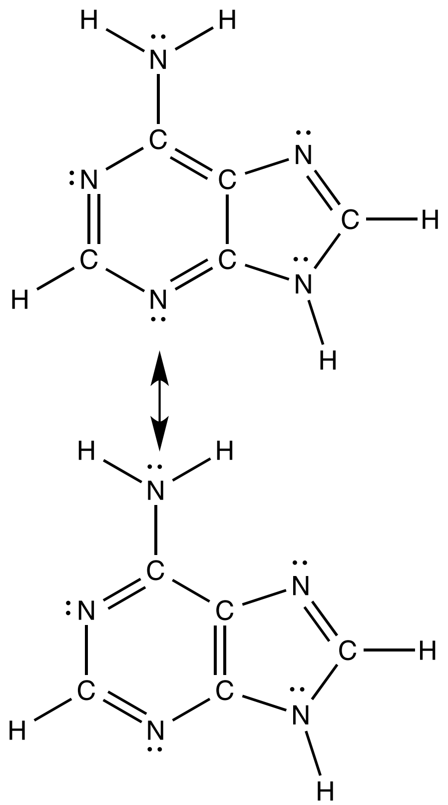
After mixing the total solution volume is (100.0 + 50.0) mL = 150.0 mL. Hence, the concentration of $Al^{3+}(aq)$ is:

concentration = number of moles / volume = n / V
= 0.0020 mol / 0.1500 L = 0.013 mol L^{-1}

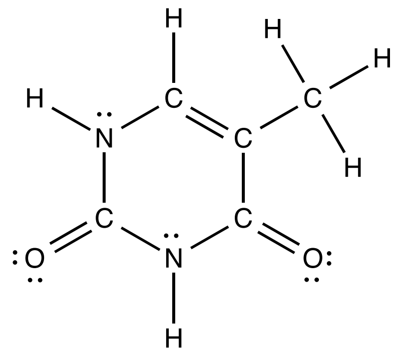
Answer: **0.013 M**

Marks
5

- By adding double bonds and lone pairs, complete the structural formulae of the nitrogen bases adenine and thymine below.

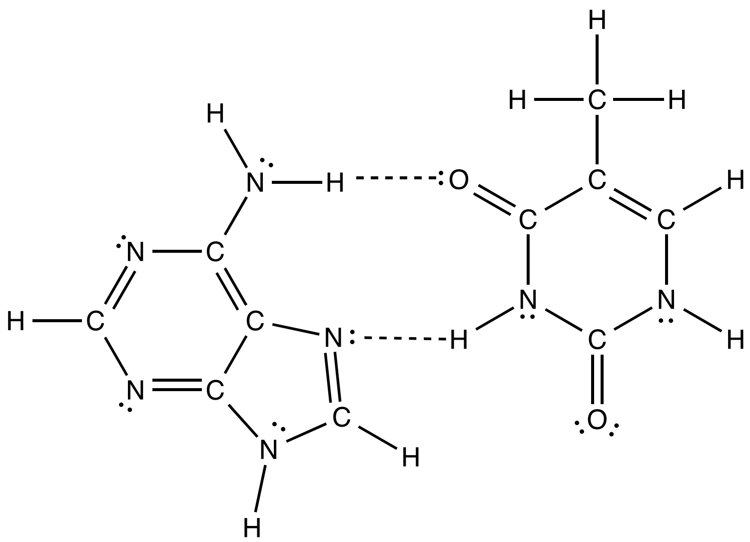


adenine



thymine

In DNA, these two molecules interact through two hydrogen bonds. Redraw the structures below showing the alignment of the two molecules that allows this to occur and clearly show the hydrogen bonds.



Marks
9

- Rechargeable nickel-cadmium batteries normally operate (discharge) with the following oxidation and reduction half-cell reactions.

$Cd(s) + 2OH^-(aq) \rightarrow Cd(OH)_2(s) + 2e^-$ $E° = 0.82$ V

$NiO(OH)(s) + H_2O(l) + e^- \rightarrow Ni(OH)_2(s) + OH^-(aq)$ $E° = 0.60$ V

Write out a balanced overall cell reaction.

$Cd(s) + 2NiO(OH)(s) + 2H_2O(l) \rightarrow Cd(OH)_2(s) + 2Ni(OH)_2(s)$

Calculate the overall cell potential.

$E°_{cell} = E°_{oxidation} + E°_{reduction} = (0.82 + 0.60)$ V = +1.42 V

Answer: **+1.42 V**

Using your balanced cell reaction, briefly explain why the cell potential does not change as the battery discharges itself.

Cell potentials depend on concentrations (as quantified using the Nernst equation). The overall cell reaction only involves solids and a pure liquid. The concentration of solids and pure liquids are constant during the reaction (until they are used up). As the concentrations do not change, the cell potential remains constant (until the battery is used up.)

Write out the balanced overall reaction that occurs when this battery is being recharged.

$Cd(OH)_2(s) + 2Ni(OH)_2(s) \rightarrow Cd(s) + 2NiO(OH)(s) + 2H_2O(l)$

A current of 2.75 A is measured during recharging with an external potential of 2.0 V. After 5.00 minutes charging, how many moles of Cd(s) will be redeposited?

The number of moles of electrons used is given by:

number of moles of e^- = It / F
$= (2.75 \text{ A}) \times (5.00 \times 60.00 \text{ s}) / (96485 \text{ C mol}^{-1})$
= 0.00855 mol

From the half-cell reaction, each mole of Cd(s) requires 2 mol of electrons. The number of moles of Cd(s) redeposited is therefore:

number of moles of Cd(s) = ½ × 0.00855 mol = 0.00428 mol

Answer: **0.00428 mol**

Marks
4

- A certain mixture of gases containing 0.24 mol of He, 0.53 mol of N_2 and 0.05 mol of Ne is placed in a container with a piston that maintains it at a total pressure of 1.0 atm. This gas mixture is now heated from its initial temperature of 290 K to 370 K by passing 2.08 kJ of energy into it.
  Calculate the volume occupied by the gas at 370 K.

The total number of moles of gas = (0.24 + 0.53 + 0.05) mol = 0.82 mol.

Using the ideal gas law, *PV* = *nRT*, the volume occupied is:

$$\begin{aligned} V &= nRT / P \\ &= (0.82 \text{ mol}) \times (0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}) \times (370 \text{ K}) / (1.0 \text{ atm}) \\ &= 25 \text{ L} \end{aligned}$$

Answer: **25 L**

Calculate the heat capacity of the gas mixture (in J K^{-1} mol^{-1}).

The temperature change = ΔT = (370 – 290) = 80 K. This is produced using a heat change of 2.08 kJ.

Using $q = nC\Delta T$, the molar heat capacity is:

$$C = q / n\Delta T = (2.08 \times 10^3 \text{ J}) / (0.82 \text{ mol} \times 80 \text{ K}) = 30 \text{ J K}^{-1} \text{ mol}^{-1}$$

Answer: **30 J K^{-1} mol^{-1}**

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Marks
8

- Nitrogen and acetylene gases react to form hydrogen cyanide according to the reaction

$$N_2(g) + C_2H_2(g) \rightleftharpoons 2HCN(g) \quad K_c = 2.3 \times 10^{-4} \text{ at } 300\ °C$$

Write out the equilibrium constant expression for K_c for this reaction as shown above.

$$K_c = \frac{[HCN(g)]^2}{[N_2(g)][C_2H_2(g)]}$$

The value of K_p for this reaction at 300 °C is also 2.3×10^{-4}. Why are the values of K_p and K_c the same for this reaction?

The number of moles of gas remains constant during the reaction. 2 mol of gas react to give 2 mol of product gas.

Write a balanced equation and calculate the value of the equilibrium constant K_c' for the formation of 1.0 mol of hydrogen cyanide gas from nitrogen and acetylene gases.

$$\tfrac{1}{2} N_2(g) + \tfrac{1}{2} C_2H_2(g) \rightleftharpoons HCN(g)$$

For this reaction, $K_c' = \dfrac{[HCN(g)]}{[N_2(g)]^{1/2}[C_2H_2(g)]^{1/2}}$

$= K_c^{1/2} = 0.015$

Answer: **0.015**

What is the equilibrium concentration of HCN(g) if nitrogen and acetylene are mixed so that both are at starting concentrations of 1.0 mol L^{-1}?

The reaction table for this is:

	½ $N_2(g)$	½ $C_2H_2(g)$	$\rightleftharpoons$	HCN(g)
initial	**1.0**	**1.0**		**0**
change	$-x$	$-x$		$+2x$
equilibrium	$1.0 - x$	$1.0 - x$		$2x$

Hence, the equilibrium constant expression in terms of x is:

$$K_c = \frac{[HCN(g)]}{[N_2(g)]^{1/2}[C_2H_2(g)]^{1/2}} = \frac{(2x)}{(1.0 - x)^{1/2}(1.0 - x)^{1/2}} = \frac{2x}{(1.0 - x)} = 0.015$$

ANSWER CONTINUES ON THE NEXT PAGE

So,

$2x = 0.015 - 0.015x$

$2.015x = 0.015$

$x = 0.0074$

ˆ

Hence,

[HCN(g) = $2x$ N = 0.015 M

(The 'small x' approximation can be used but as no quadratic needs to be solved, this is unnecessary.)

Answer: **0.015 M**

Marks
3

- The boiling point of NH_3 is –33 °C and that of HF is +20 °C. Explain this difference in terms of the strengths of the intermolecular forces between these molecules.

The strongest intermolecular force in both comes from hydrogen bonding.

Each HF molecule possesses 3 lone pairs on F and 1 H. HF molecules on average make 2 H-bonds.

Each NH_3 molecule possesses 1 lone pair on N and 3 H. NH_3 molecules on average also make 2 H-bonds.

As fluorine is more electronegative than nitrogen, the H-F bonds are much more polar than the N-H bonds. Due to the higher partial charges on H and F in HF, a hydrogen bond between HF molecules is stronger than that between NH_3 molecules.

The higher boiling point of HF is thus due to stronger H-bonds.

Explain why the boiling point of water (100 °C) is higher than both HF and NH_3.

The polarity of the O-H bonds in H_2O is intermediate between that of H-F and N-H bonds. It is expected that the individual H-bonds between H_2O molecules will also be intermediate in strength.

Each H_2O molecule possesses 2 lone pairs on O and 2 H. H_2O molecules are thus able to form an average of 4 H-bonds.

H_2O has a higher boiling point than NH_3 because (i) the H-bonds are stronger and (ii) it contains twice as many H-bonds.

H_2O has a higher boiling point than HF because it contains twice as many H-bonds, despite these being individually weaker.

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Marks

- Write the equation whose enthalpy change represents the standard enthalpy of formation of $NO(g)$. **3**

½ $N_2(g)$ + ½ $O_2(g)$ → $NO(g)$

Given the following data, calculate the standard enthalpy of formation of $NO(g)$.

$$N_2(g) + 2O_2(g) \rightleftharpoons 2NO_2(g) \qquad \Delta H° = 66.6 \text{ kJ mol}^{-1}$$

$$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g) \qquad \Delta H° = -114.1 \text{ kJ mol}^{-1}$$

Using $\Delta_{rxn}H° = \Sigma m\Delta_f H°$(products) - $\Sigma n\Delta_f H°$(reactants), the enthalpy of these two reactions are:

(1) $\Delta H° = [2\Delta_f H°(NO_2(g))] = 66.6 \text{ kJ mol}^{-1}$

(2) $\Delta H° = [2\Delta_f H°(NO_2(g))] - [2\Delta_f H°(NO(g))] = -114.1 \text{ kJ mol}^{-1}$

where $\Delta_f H°(N_2(g)) = \Delta_f H°(O_2(g)) = 0$ has been used for elements already in their standard states.

Substituting (1) into (2) gives:

$(66.6 \text{ kJ mol}^{-1}) - [2\Delta_f H°(NO(g))] = -114.1 \text{ kJ mol}^{-1}$

$\Delta_f H°(NO(g))] = 90.4 \text{ kJ mol}^{-1}$

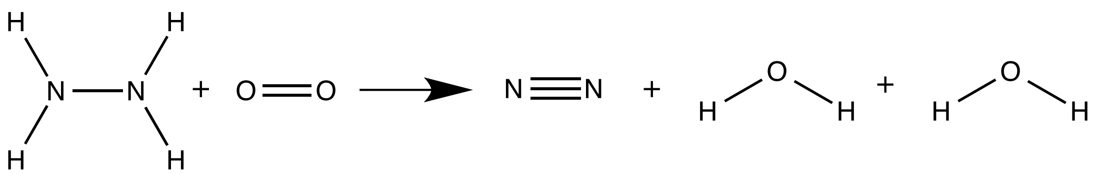
Answer: **+90.4 kJ mol^{-1}**

- Hydrazine, N_2H_4, burns completely in oxygen to form $N_2(g)$ and $H_2O(g)$. Use the bond enthalpies given below to estimate the enthalpy change for this process. **3**

Bond	Bond enthalpy (kJ mol^{-1})	Bond	Bond enthalpy (kJ mol^{-1})
N–H	391	O=O	498
N–N	158	O–O	144
N=N	470	O–H	463
N≡N	945	N–O	214

The chemical equation for the combustion is:

$N_2H_4(g) + O_2(g) \rightarrow N_2(g) + 2H_2O(g)$



ANSWER CONTINUES ON THE NEXT PAGE

This requires:

- **breaking 1 N-N bond, 4 N-H bonds and 1 O=O bond**
- **making 1 N≡N bond and 4 O-H bonds**

Breaking bonds is endothermic:

$\Delta_{breaking}H° = (158 + 4 \times 391 + 498)$ kJ mol^{-1} = +2220 kJ mol^{-1}

Making bonds is exothermic:

$\Delta_{making}H° = -(945 + 4 \times 463)$ kJ mol^{-1} = -2797 kJ mol^{-1}

Overall,

$\Delta H° = \Delta_{breaking}H° + \Delta_{making}H° = (2220 – 2797)$ kJ mol^{-1} = -577 kJ mol^{-1}

Answer: **-577 kJ mol^{-1}**