Topics in the June 2012 Exam Paper for CHEM1001

Click on the links for resources on each topic.

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- Molecules and Ions
- Elements and Atoms
- Chemical Equations
- Stoichiometry

2012-J-3:

- Lewis Model of Bonding
- VSEPR

2012-J-4:

- Stoichiometry
- Elements and Atoms

2012-J-5:

- Atomic Energy Levels
- Lewis Model of Bonding
- VSEPR

2012-J-6:

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2012-J-7:

- Molecules and Ions
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2012-J-8:

- Introduction to Electrochemistry
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- Chemical Equilibrium

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- Types of Intermolecular Forces

2012-J-12:

- Thermochemistry

- First Law of Thermodynamics

Marks

- Complete the following table. **2**

Name	Formula
calcium nitride	**Ca_3N_2**
carbon tetrabromide	**CBr_4**
iron(III) oxide	Fe_2O_3
sulfuric acid	**H_2SO_4**

- Explain why relative atomic masses are not always close to an integer. For example, copper has a reported value of 63.54. **2**

Many elements consist of isotopes, *i.e.* atoms with different numbers of neutrons and hence different atomic masses. The atomic mass of each isotope is close to an integer value. The relative atomic mass of an element is calculated using all these different isotopic masses and their relative percentages.

- Analysis of a black-coloured mineral called pitchblende returned the following percentage composition by weight: 84.80% uranium and 15.20% oxygen. What is the empirical formula of this compound? **2**

The mineral contains 84.80% U and 15.20% O.

	U	**O**
percentage	**84.80**	**15.20**
divide by atomic mass	$\frac{84.80}{238.03} = 0.356$	$\frac{15.20}{16.00} = 0.950$
divide by smallest value	**1**	**2.67**

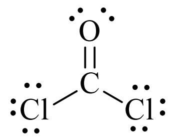
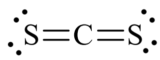
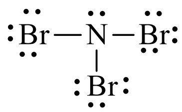
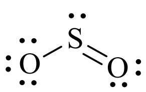
The ratio of U : O is 1 : 2.67. The simplest whole number ratio can be obtained by multiplying this by 3 to give U : O equal to 3 : 8.

The empirical formula is U_3O_8.

Answer: **U_3O_8**

Marks
6

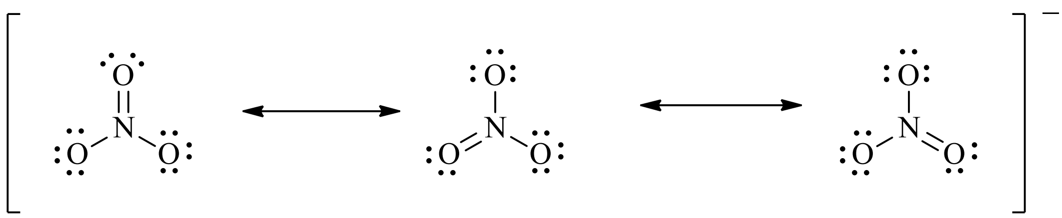
- Complete the following table, including resonance structures where appropriate. The central atom is underlined.

Species	Lewis structure(s)	Is the molecule polar?
<u>C</u>OCl_2		**yes**
<u>C</u>S_2		**no**
<u>N</u>Br_3		**yes**
<u>S</u>O_2		**yes**

- What is resonance? Give at least one example.

2

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.



For example, in NO_3^-, 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.

Marks

- Balance the following equation: **3**

$$NH_3(g) + O_2(g) \rightarrow NO(g) + H_2O(l)$$

$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$

Calculate the mass of NH_3 required to produce 140. g of water.

The molar mass of H_2O is:

molar mass = [2 × 1.008 (H) + 16.00 (O)] g mol^{-1} = 18.016 g mol^{-1}

Hence, the number of moles of water produced is:

number of moles = mass / molar mass = (140. g) / (18.016 g mol^{-1}) = 7.771 mol

From the balanced equation, 4 mol of NH_3 will produce 6 mol of H_2O. Hence, to produce 7.771 mol of H_2O so:

number of moles of NH_3 = (4/6) × 7.771 mol = 5.18 mol

The molar mass of NH_3 is:

molar mass = [14.01 (N) + 3 × 1.008 (H)] g mol^{-1} = 17.034 g mol^{-1}

The mass of NH_3 in 5.18 mol is therefore:

mass = number of moles × molar mass = (5.18 mol) × (17.034 g mol^{-1}) = 88.2 g

Answer: **88.2 g**

- Describe Rutherford's experiment that showed atoms consisted of a concentrated positive charge with a high mass. Make sure you discuss the observations and the conclusions drawn. **2**

A stream of positively charged alpha particles was fired at a thin sheet of gold foil. Most of the particles passed straight through or were slightly deflected, but the occasional one was reflected back towards the source.

The conclusion drawn was that atoms consist of mostly empty space with a small, dense, positively charged nucleus.

Marks

- For a single atom, complete the following table. If more than one quantum number is possible, give all correct possibilities.

6

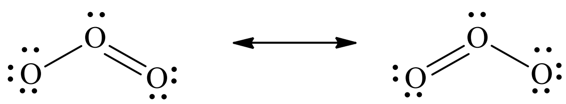
Name	Maximum number of electrons contained	Quantum numbers	
		n	l
1*s* orbital	**2**	**1**	**0**
2p_x orbital	**2**	**2**	**1**
3*d* subshell	**10**	**3**	**2**
2nd shell	**8**	**2**	**0 and 1**

- The σ-bonding in two plausible structures of ozone, O_3, is shown below. Complete each structure by adding electrons and/or π-bonds as appropriate.

3

Which of these geometries does ozone adopt? Give reasons for your answer.

Ozone adopts the non-cyclic structure. The cyclic structure is very strained with bond angles of 60° instead of 109.5°, making it very unstable. In contrast, the second structure is stabilised by resonance.



Ozone does not contain 1 double and 1 single bond. Both the O-O bonds are exactly the same length and true structure is a sort of average of the two Lewis structures shown. The energy of the true structure is lower than the theoretical energy for either of the given structures. This energy difference is known as resonance stabilisation energy.

Marks

- Describe the differences between a 1*s* atomic orbital and a 2*s* atomic orbital. **2**

1s orbital is smaller, closer to the nucleus, has a lower energy and has 0 nodes.
2s orbital is larger, further from the nucleus, has a higher energy and has 1 node.

Both orbitals are spherical.

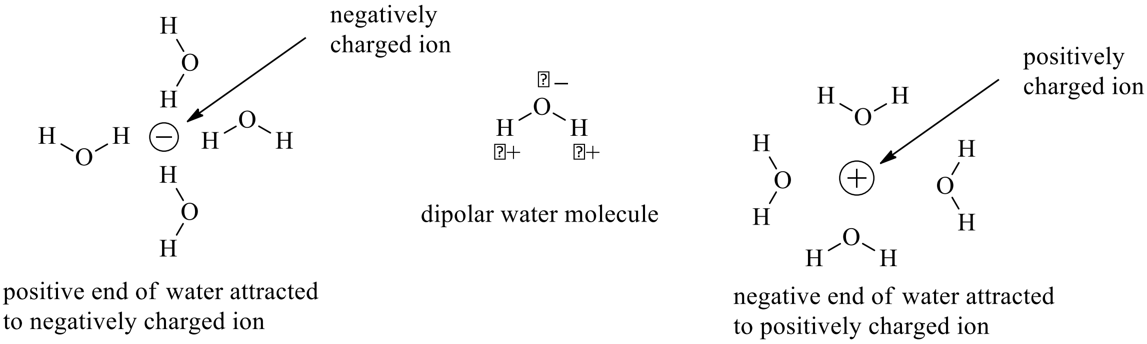
- Complete the following table. **3**

Species	Full electron configuration
gallium atom	$\mathbf{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^1}$
P^{3-}	$\mathbf{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6}$
K^+	$\mathbf{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6}$

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY

• Depict the arrangement of water molecules around an ion. Explain why many ionic compounds are soluble in water.

Marks
3



Water is a dipolar molecule. The positive ends (H) can directly interact with negative ions, while the negative end (O) can directly interact with positive ions. These interactions (called hydration enthalpy) are often sufficient to overcome the lattice enthalpy (the energy required to break up the ionic solid into its constituent ions).

• The equation for the detonation of nitroglycerine, $C_3H_5N_3O_9(l)$, is given below. **3**

$$4C_3H_5N_3O_9(l) \rightarrow 6N_2(g) + 12CO_2(g) + 10H_2O(g) + O_2(g)$$

What mass of nitroglycerine is required to produce 720 L of product gases at 1800 °C and 1.00 atm? Assume all gases behave as ideal gases. Show all working.

Using the ideal gas equation, $PV = nRT$, the total number of moles of gas produced is:

$$n = PV / RT = (1.00 \text{ atm} \times 720 \text{ L}) / (0.08206 \text{ atm L K}^{-1} \text{ mol}^{-1} \times (1800 + 273) \text{ K}) = 4.23 \text{ mol}$$

From the chemical equation, detonation of 4 mol of nitroglycerine gives (6 + 12 + 10 + 1) mol = 29 mol of gases. Therefore:

number of moles of nitroglycerine required = (4/29) × 4.23 mol = 0.584 mol

The molar mass of nitroglycerine is:

$$\text{molar mass} = [3 \times 12.01 + 5 \times 1.008 \text{ (H)} + 3 \times 14.01 \text{ (N)} + 9 \times 16.00 \text{ (O)}] \text{ g mol}^{-1} = 227.1 \text{ g mol}^{-1}$$

The mass of nitroglycerine is therefore:

$$\text{mass} = \text{number of moles} \times \text{molar mass} = (0.582 \text{ mol}) \times (227.1 \text{ g mol}^{-1}) = 130 \text{ g}$$

Answer: **130 g**

Marks 8

- A galvanic cell has the following cell reaction:

  $D(s) + 2Zn^{2+}(aq) \rightarrow 2Zn(s) + D^{4+}(aq)$ $E° = 0.18$ V

  Write the overall cell reaction in shorthand cell notation.

$D(s) \mid D^{4+}(aq) \parallel Zn^{2+}(aq) \mid Zn(s)$

Is the reaction spontaneous? Why?

Yes. $E°$ is positive and hence $\Delta G°$ is negative.

Which electrode is the anode? **D**

Write the equation for the half-reaction that occurs at the anode?

$D(s) \rightarrow D^{4+}(aq) + 4e^-$

What is the standard reduction potential of the D^{4+}/D redox couple?

The two half equations are:

$D(s) \rightarrow D^{4+}(aq) + 4e^-$ $E°_{oxidation}(D^{4+}/D)$
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $E°_{reduction}(Zn^{2+}/Zn)$

where $E°_{oxidation}(D^{4+}/D) + E°_{reduction}(Zn^{2+}/Zn) = 0.18$ V. From the standard reduction potential for $Zn^{2+}(aq)$ is -0.76 V. Therefore:

$E°_{oxidation}(D^{4+}/D) = (0.18 - (-0.76))$ V $= 0.94$ V

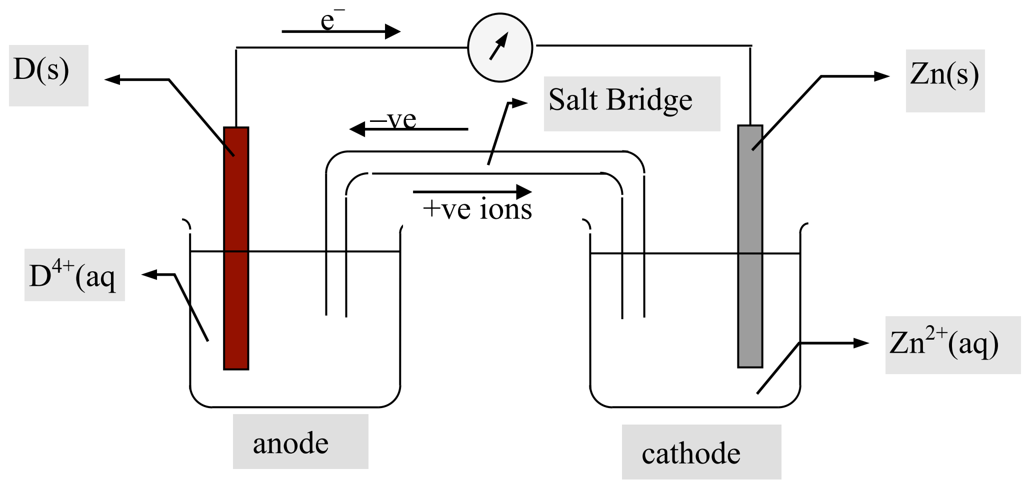
The standard reduction potential for the D^{4+}/D redox couple is the reverse of this:

$E°_{reduction}(D^{4+}/D) = -0.94$ V

Answer: **-0.94 V**

ANSWER CONTINUES OVER THE PAGE

Draw, labelling all essential components, a cell diagram for this cell.



Marks

- A 120.0 g piece of copper is heated to 80.0 °C before being added to 150.0 mL of water at 25.0 °C. What is the final temperature of the mixture? The specific heat capacity of copper is 0.385 J g^{-1} K^{-1} and the specific heat capacity of water is 4.18 J g^{-1} K^{-1}. **3**

The copper will cool down and the water will heat up when the two are mixed. The final temperature, T_f, will be the same for both.

For the copper,

$$q_{copper} = m\,c\,\Delta T = (120.0 \text{ g}) \times (0.385 \text{ J g}^{-1} \text{ K}^{-1}) \times \Delta T_{copper}$$

For the water,

$$q_{water} = m\,c\,\Delta T = (150.0 \text{ g}) \times (4.18 \text{ J g}^{-1} \text{ K}^{-1}) \times \Delta T_{water}$$

As the heat lost by the copper is gained by the water, $q_{water} = -q_{copper}$:

$$(150.0 \text{ g}) \times (4.18 \text{ J g}^{-1} \text{ K}^{-1}) \times \Delta T_{water} = -\,(120.0 \text{ g}) \times (0.385 \text{ J g}^{-1} \text{ K}^{-1}) \times \Delta T_{copper}$$

or

$$627 \times \Delta T_{water} = -\,46.2 \times \Delta T_{copper}$$

Using $\Delta T_{water} = (T_f - 25.0)$ °C and $\Delta T_{copper} = (T_f - 80.0)$ °C gives:

$$T_f = 28.8 \text{ °C}$$

Answer: **28.8 °C**

- H^+ is reduced to H_2 in an electrochemical cell. What is the total charge transferred when a current of 2 A is passed through the cell for 20 minutes? **5**

Total charge = current × time = (2 A) × (20 × 60 s) = 2400 C = 2000 C (to 1 s.f.)

Answer: **2000 C**

What amount of H_2 (in mol) is produced under these conditions?

The equation for the reduction of H^+ to H_2 is:

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

The number of moles of electrons in 2000 C is:

moles of electrons = total charge / Faraday's constant = 2000 C / 96485 C mol^{-1} = 0.02 mol

ANSWER CONTINUES ON THE NEXT PAGE

Two moles of electrons is required for each mole of H_2. Hence the number of moles of H_2 is half of this:

number of moles of H_2 = (1/2) × 0.02 mol = 0.01 mol

Answer: **0.01 mol**

What volume would this gas occupy at 25 °C and 90 kPa?

As 1 atm = 101.3 kPa, 90 kPa corresponds to:

pressure = (90 / 101.3) atm = 0.89 atm

Using the ideal gas equation, $PV = nRT$:

$$V = nRT / P$$
$$= (0.01 \text{ mol}) \times (0.08206 \text{ atm L K}^{-1} \text{ mol}^{-1}) \times ((25 + 273) \text{ K}) / (0.89 \text{ atm})$$
$$= 0.3 \text{ L}$$

Answer: **0.3 L**

Marks 4

- Consider the following reaction.

$$SO_2(g) + NO_2(g) \rightleftharpoons SO_3(g) + NO(g)$$

An equilibrium mixture in a 1.00 L vessel was found to contain $[SO_2(g)] = 0.800$ M, $[NO_2(g)] = 0.100$ M, $[SO_3(g)] = 0.600$ M and $[NO(g)] = 0.400$ M. If the volume and temperature are kept constant, what amount of NO(g) needs to be added to the reaction vessel to give an equilibrium concentration of $NO_2(g)$ of 0.300 M?

From the chemical equation,

$$K_{eq} = \frac{[SO_3(g)][NO(g)]}{[SO_2(g)][NO_2(g)]}$$

As the original mixture is at equilibrium:

$$K_{eq} = \frac{[SO_3(g)][NO(g)]}{[SO_2(g)][NO_2(g)]} = \frac{(0.600)(0.400)}{(0.800)(0.100)} = 3.00$$

This equilibrium is now disturbed by the addition of *x* M of NO(g). To re-establish equilibrium, the reaction will shift to the left by an unknown amount *y*. The reaction table for this is:

	$SO_2(g)$	$NO_2(g)$		$SO_3(g)$	$NO(g)$
initial	**0.800**	**0.100**	$\rightleftharpoons$	**0.600**	**0.400 + x**
change	**+y**	**+y**		**-y**	**-y**
equilibrium	**0.800 + y**	**0.100 + y**		**0.600 – y**	**0.400 + x - y**

As $[NO_2(g)] = 0.300$ M at the new equilibrium, $y = (0.300 – 0.100)$ M = 0.200 M. Hence, the new equilibrium concentrations are:

$[SO_2(g)] = (0.800 + 0.200)$ M = 1.000 M
$[NO_2(g)] = 0.300$ M
$[SO_3(g)] = (0.600 – 0.200)$ M = 0.400 M
$[NO(g)] = (0.400 + x – 0.200)$ M = $(0.200 + x)$ M

As the system is at equilibrium,

$$K_{eq} = \frac{[SO_3(g)][NO(g)]}{[SO_2(g)][NO_2(g)]} = \frac{(0.400)(0.200+x)}{(1.000)(0.300)} = 3.00$$

Solving this gives $x = 2.05$ M. As the reaction is carried out in a 1.00 L container, this is also the number of moles required.

Answer: **2.05 mol**

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY

Marks

- Rank the following compounds in order of increasing boiling point? Justify your answer. **3**

$CH_3CH_2OCH_2CH_3$, CH_3OH, CH_4, CH_3CH_3, CH_3CH_2OH

CH_4 < CH_3CH_3 < $CH_3CH_2OCH_2CH_3$ < CH_3OH < CH_3CH_2OH

Only weak dispersion forces act in CH_4 and CH_3CH_3. The bigger molecule has more interactions and hence the higher b.p. $CH_3CH_2OCH_2CH_3$ is a bigger molecule than CH_4 and CH_3CH_3, so has more dispersion forces. It also has dipole-dipole forces due to the polarised C-O bonds.

CH_3OH and CH_3CH_2OH have hydrogen bonds due to the very electronegative O atom bonded to the H atom. These H-bonds are much stronger than the dispersion and dipole-dipole forces in the other compounds and hence these two compounds have the highest boiling points. CH_3CH_2OH has more dispersion forces than CH_3OH, so it has the highest boiling point.

- Melting points of the hydrogen halides increase in the order HCl < HBr < HF < HI. Explain this trend. **2**

There are two competing intermolecular forces at play:

- **Dipole-dipole forces increase as the halogen becomes more electronegative (I < Br < Cl < F).**
- **Dispersion forces are dependent on the polarisability of the atoms and increase with the size of the halogen.**

Dispersion force dominate in HCl, HBr and HI and determines the order of their melting points.

The dipole-dipole force in HF is so strong (due to the very small and very electronegative F atom) that it is given a special name - a hydrogen bond. This causes HF to have an anomalously high melting point, which just happens to lie between that of HBr and HI.

Marks 6

- Use average bond dissociation enthalpies given below to calculate the molar enthalpy change for the following chemical transformation:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Bond	H–H	N–H	N≡N
ΔH / kJ mol^{-1}	436	391	945

The reaction requires:

- **breaking 1 N≡N and 3 H-H bonds and**
- **formation of 6 N-H bonds**

The enthalpy required to break these bonds is:

$\boldsymbol{\Delta H}$ **(bond breaking) = [945 (N≡N) + 3 × 436 (H-H)] kJ mol^{-1}**
= +2253 kJ mol^{-1}

Enthalpy is *released* by making the new bonds is:

$\boldsymbol{\Delta H}$ **(bond making) = - [6 × 391 (N-H)] kJ mol^{-1}**
= -2346 kJ mol^{-1}

The overall enthalpy change is therefore:

$\boldsymbol{\Delta H}$ **= [(+2253) + (-2346)] kJ mol^{-1} = -93 kJ mol^{-1}**

Answer: **-93 kJ mol^{-1}**

What is the standard enthalpy of formation, $\Delta_f H°$, of NH_3(g)? **-47 kJ mol^{-1}**

(Note: the reaction in the question produces 2 mol of NH_3 so the enthalpy of formation is half of the enthalpy change of this reaction.)

The standard enthalpy of formation of hydrazine, N_2H_4(g) is +96 kJ mol^{-1}. Calculate the strength of the N–N single bond in hydrazine.

The standard enthalpy of formation of N_2H_4(g) corresponds to the reaction:

$$N_2(g) + 2H_2(g) \rightarrow N_2H_4(g)$$

This reaction requires:

- **breaking 1 N≡N and 2 H-H bonds and**
- **formation of 4 N-H and 1 N-N bond**

The enthalpy required to break these bonds is:

$\boldsymbol{\Delta H}$ **(bond breaking) = [945 (N≡N) + 2 × 436 (H-H)] kJ mol^{-1}**
= +1817 kJ mol^{-1}

ANSWER CONTINUES ON THE NEXT PAGE

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- **formation of 4 N-H and 1 N-N bond**

The enthalpy required to break these bonds is:

$$\Delta H \text{ (bond breaking)} = [945 \text{ (N≡N)} + 2 \times 436 \text{ (H-H)}] \text{ kJ mol}^{-1}$$
$$= +1817 \text{ kJ mol}^{-1}$$

Enthalpy is *released* by making the new bonds is:

$$\Delta H \text{ (bond making)} = -[4 \times 391 \text{ (N-H)} + x \text{ (N-N)}] \text{ kJ mol}^{-1}$$
$$= -(1564 + x) \text{ kJ mol}^{-1}$$

The overall enthalpy change is equal to the enthalpy of formation of N_2H_4:

$$\Delta_f H = [(+1817) + -(1564 + x)] \text{ kJ mol}^{-1} = +96 \text{ kJ mol}^{-1}$$

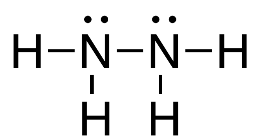
So

$$x = +157 \text{ kJ mol}^{-1}$$

Answer: **+157 kJ mol^{-1}**

Suggest why the N–N single bond in hydrazine is much weaker than the N–H and H–H bonds. Hint: Draw its Lewis structure.

Each N atom in hydrazine has a lone pair of electrons and is the negative end of a dipole formed with the H atoms.



These lone pairs repel each other, weakening the bond.