Name: ____________________ **Student number:** ______________

Chemistry 1E03 **Test 1** **Oct. 18, 2013**

McMaster University ***VERSION 1 SOLUTIONS***

1. The amount of SO_2 present in a sample of air can be determined by titration with a solution of $KMnO_4$**.** **How many moles** of SO_2 are present in a sample of air, if the sample of SO_2 is consumed by 5.87 mL of 0.00800 M $KMnO_4$ solution?

$$5\ SO_2 + 2\ MnO_4^{-} + 2\ H_2O \longrightarrow 5\ SO_4^{2-} + 2\ Mn^{2+} + 4\ H^{+}$$

.

A) 1.88×10^{-5}
B) 9.39×10^{-6}
C) 5.70×10^{-5}
D) 9.79×10^{-4}
E) 1.17×10^{-4}

moles of MnO_4^- = (0.00587 L)(0.00800 mol.L^{-1}) = 4.70×10^{-5} mol

moles of SO_2 = (5/2)*(# moles of MnO_4^-) = 2.5* 4.70×10^{-5} mol = 1.17×10^{-4} mol

Note that the K^+ cation is a spectator in this reaction and can be ignored.

2. The cation $^{33}\mathbf{S}^{+}$ contains

.

A) 17 neutrons, 16 protons, 17 electrons
B) 33 neutrons, 17 protons, 16 electrons
C) 16 neutrons, 16 protons, 15 electrons
D) 17 neutrons, 17 protons, 16 electrons
E) 17 neutrons, 16 protons, 15 electrons

Sulfur has atomic number 16. This means it has 16 protons. The left superscript, 33, is the mass number. This equals the total number of protons and neutrons. Subtracting the number of protons gives 33 – 16 = 17 neutrons. Since this is a 1+ cation, there are fewer electrons (by 1) than protons: 16 – 1 = 15 electrons.

3. A certain gas weighs 0.988 g and occupies a volume of 1.00 L at a pressure of 0.928 atm and a temperature of 46.8°C. Which gas could this be?

.

A) CH_4
B) CF_4
C) CO
D) CO_2
E) Cl_2

PV = nRT

$$n = PV/RT = \frac{(0.928\text{ atm})(1.00\text{ L})}{(0.082\text{ L.atm.K}^{-1}mol^{-1})(273.16+46.8)K} = 0.0354mol$$

MW = m/n = (0.988g)/(0.0354mol)=27.9 g/mol ~ 28 g/mol - this is MW of CO

4. What is the empirical formula for iron(III) sulfide?

.

A) Fe_3S_2
B) FeS
C) Fe_2S
D) Fe_2S_3
E) FeS_3

Fe(III) indicates the Fe has a formal oxidation state of 3+
Sulfide indicates the sulfur has a formal oxidation state of 2-

To have an overall neutral compound, there must be 3 S and 2 Fe

5. What **mass** (in g) of **CaO(s)** is produced when 6.80 g of calcium metal reacts with 2.00 L of oxygen at 298 K and 1.00 atm pressure?
   .
   A) 6.89
   B) 9.17
   C) 10.1
   D) 9.87
   E) 8.14

The balanced reaction is

$$Ca(s) + \tfrac{1}{2} O_2(g) \rightarrow CaO(s)$$

6.80 g of Ca corresponds to

$$6.80 \text{ g} / 40.078 \text{ g mol}^{-1} = 0.169_7 \text{ mol.}$$

2.00 L of O_2 gas corresponds to

$$n = PV/(RT) = 1.00 \text{ atm} \times 2.00 \text{ L} /(0.08206 \text{ L atm K}^{-1}\text{mol}^{-1} \times 298 \text{ K})$$
$$= 0.0817_9 \text{ mol}$$

To consume all of the oxygen requires $2 \times 0.0817_9$ mol = 0.1636 mol of calcium. Since more calcium is available, the oxygen will be consumed – oxygen is the limiting reactant. Therefore,

$2 \times 0.0817_9$ mol = 0.1636 mol of CaO is produced

This corresponds to 0.1636 mol × (40.078+15.9994) g mol^{-1} = 9.17 g.

6. Consider the following unbalanced redox reaction in **acidic solution**. What is the **coefficient** of H^+, and the **number of electrons** transferred, when the reaction is **balanced** using the **smallest whole-number coefficients**?

$$H_2S + MnO_4^- \longrightarrow Mn^{2+} + SO_4^{2-}$$

   .
   A) 6, 80
   B) 14, 40
   C) 12, 80
   D) 4, 80
   E) 8, 40

Assign ONs:

$H_2S + MnO_4^- \longrightarrow Mn^{2+} + SO_4^{2-}$

-2 +7 +2 +6

S goes from -2 to +6 – i.e., each S atom loses 8 electrons. Mn goes from +7 to +2 – each Mn atom gains 5 electrons. The (unbalanced – with respect to oxygen and hydrogen) half reactions are therefore

$$H_2S \longrightarrow SO_4^{2-} + 8\,e^-$$

and

$MnO_4^- + 5\ e^- \rightarrow Mn^{2+}$

To balance electrons we must multiply the first equation by 5 and the second by 8. Adding the resulting half reactions gives the following reaction:

$5\ H_2S + 8\ MnO_4^- + 40\ e^- \rightarrow 8\ Mn^{2+} + 5\ SO_4^{2-} + 40\ e^-$

Here the electrons cancel. However, in the above form, we can see that 40 electrons are transferred.

We now balance O and H atoms, noting that the solution is acidic, starting with O. There are 8×4 = 32 O atoms on the left, and 5×4 = 20 O atoms on the right. Add 12 H_2O to the right to balance the O atoms.

7. You are given two solutions, A and B. One contains **KCl(aq)**, while the other contains **$NaNO_3$(aq)**. Which **one** of the following test solutions can be used to distinguish which of solutions A and B is KCl(aq)?

.

A) H_2SO_4(aq)
B) $KMnO_4$(aq)
C) $Ba(OH)_2$(aq)
D) $CuClO_4$(aq)
E) $AgCH_3COO$(aq)

The solutions listed A through E are called "test solutions". These are solutions added to A and B in order to distinguish which one is KCl(aq). If $AgCH_3COO$(aq) is added to A and B, the solution (A or B) with KCl(aq) will precipitate AgCl(s), as AgCl is insoluble. The other solution which contains $NaNO_3$(aq) will not exhibit any reaction – $AgNO_3$ and $NaCH_3COO$ are soluble. Adding any of the other test solutions produces no precipitate or other visible indication of either KCl(aq) or $NaNO_3$(aq).

8. Which **one** reaction is **not** an **acid-base** reaction?

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A) $NaH_2PO_4(aq) + Li_2CO_3(aq) \rightarrow Na^+(aq) + HPO_4^{2-}(aq) + 2\ Li^+(aq) + HCO_3^-(aq)$
B) $BaO(s) + 2\ HCl(aq) \rightarrow BaCl_2(aq) + H_2O(l)$
C) $KHCO_3(aq) + KOH(aq) \rightarrow K_2CO_3(aq) + H_2O(l)$
D) $3\ Cu(s) + 2\ MnO_4^-(aq) + 4\ H_2O(l) \rightarrow 3\ Cu^{2+}(aq) + 2\ MnO_2(s) + 8\ OH^-(aq)$
E) $CH_3COOH(aq) + H_2O(l) \rightarrow CH_3COO^-(aq) + H_3O^+(aq)$

A, C, D and E are all acid base reactions – H^+ is transferred in each case. In A, H^+ is transferred from CH_3COOH to H_2O. In C, H^+ is transferred from $H_2PO_4^-$ to CO_3^{2-}. In D, H^+ is transferred from HCO_3^- to OH^-. In E, H^+ is transferred from H_3O^+ (this is present in an solution of strong acid, HCl) to O^{2-} (present in the ionic solid, BaO). B is a redox reaction. It has a more complex set of steps – it is not a simple H^+ transfer reaction.

9. Identify the **oxidizing agent** in the following reaction:

$$HS^-(aq) + 4\ ClO^-(aq) + OH^-(aq) \longrightarrow SO_4^{2-}(aq) + 4\ Cl^-(aq) + H_2O(l)$$

.

A) ClO^-
B) OH^-
C) Cl^-
D) HS^-
E) SO_4^{2-}

Assign oxidation numbers (ONs) – note that the ONs of all H's and O's are +1 and -2, respectively:

$HS^-(aq) + 4\ ClO^-(aq) + OH^-(aq) \longrightarrow SO_4^{2-}(aq) + 4\ Cl^-(aq) + H_2O(l)$
-2 (for S) +1 (for Cl) +6 (for S) -1 (for Cl)

We see that the ON of S increases from -2 to +6 – i.e., S is oxidized. This means $HS^-(aq)$ is the reducing agent – it reduces something else, and gets oxidized in the process.
The ON of Cl decreases from +1 to -1. Thus, $ClO^-(aq)$ is the oxidizing agent – it oxidizes HS^- and gets reduced in the process.

10. Which of the following reactions is **not** an oxidation-reduction?

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A) $Na(s) + H_2O(l) \longrightarrow Na^+(aq) + OH^-(aq)$) $+ H_2(g)$ (note missing hydrogen gas product)
0 +1 +1 0
B) $CH_4(g) + 2\ O_2(g) \longrightarrow CO_2(g) + 2\ H_2O(l)$
-4 0 +4 -2 -2
C) $Cl_2(aq) + 2\ Br^-(aq) \longrightarrow 2\ Cl^-(aq) + Br_2(aq)$
0 -1 -1 0
D) $NH_4NO_3(s) \longrightarrow N_2(g) + 2\ H_2O(l) + ½\ O_2(g)$
-3 +5 -2 0 0
E) $NH_3(aq) + HF(aq) \longrightarrow NH_4^+(aq) + F^-(aq)$
-3 -1 -3 -1 ← ONs of N and F

ONs of all atoms in choice B do not change – ONs of N and F are shown above. In every other reaction the ON of one atom decreases, while another increases (as shown).

11. A reaction is carried out in aqueous solution, in a constant pressure ("coffee-cup") calorimeter. The **temperature** of the solution is observed to **decrease**. Which **one** of the following statements **must** be **TRUE**?
.
A) The reaction is a neutralization reaction.
B) The reaction is exothermic.
C) The reaction is endothermic.
D) Work is done on the surroundings, as the reaction proceeds.
E) The reaction is a formation reaction.

Temperature decreases because heat is consumed by the reaction – it is endothermic. There is no indication whether the reaction is a formation reaction, or a neutralization reaction, and there is no indication about any volume change – i.e., whether work is done on or by the system, or neither.

12. Identify the **TRUE** statement(s):
(i) If a reaction produces a net increase in moles of gas, the work done on the system (the reaction mixture) is negative.
(ii) The melting of ice is endothermic.
(iii) The standard enthalpy of formation of $Cl_2(g)$ is zero.
.
A) i, ii, iii
B) i, ii
C) i
D) ii
E) ii, iii

i is true. A net increase in moles of gas means volume increases. This means the system is doing work on the surroundings. $w = -P\,\Delta V < 0$ because P and ΔV are positive. w is work done on the system. It is negative when work is done by the system, as it is in this case.

ii is true. Melting requires the breaking of intermolecular bonds which otherwise hold molecules or atoms in the lattice positions of the solid form.

iii is true. $Cl_2(g)$ is the standard state of the element chlorine.

13. How much **heat flow**, in kJ, is required to convert 63.0 g of ice (solid water) at −15.0°C to liquid water at 10.0°C? The enthalpy of fusion (melting) of ice at 0°C is 6.01 kJ mol^{-1}. The specific heat capacities of ice and liquid water are on the data page.
.
A) 5.72
B) 175
C) 25.6
D) 33.0
E) 15.1

(1) The ice must first be heated from −15.0°C to 0.0°C. (2) The ice is melted at 0.0°C. (3) The liquid water formed is heated from 0.0°C to 10.0°C.

Step 1 requires heat, $q_1 = C_{ice}\,\Delta T = m_{ice}\,c_{ice}\,\Delta T = 63.0\text{ g} \times 2.03\text{ J / g·°C} \times (0.0 - (-15.0))$
$= 63.0\text{ g} \times 2.03\text{ J / g·°C} \times 15.0 = 1920\text{ J} = 1.92\text{ kJ}$.

63.0 g of H_2O corresponds to 63.0 g / (18.015 g mol^{-1}) = 3.50 mol
Step 2 requires heat, $q_2 = n_{ice}\,\Delta H_{fus} = 3.50\text{ mol} \times 6.01\text{ kJ mol}^{-1}$
$= 21.0\text{ kJ}$.
Note the mass of liquid water formed equals the mass of ice we started with.
Step 3 requires heat, $q_3 = C_{liq}\,\Delta T = m_{liq}\,c_{liq}\,\Delta T = 63.0\text{ g} \times 4.18\text{ J / g·°C} \times (10.0 - 0.0)$
$= 2630\text{ J} = 2.63\text{ kJ}$.

Altogether, the heat required is
$q = q_1 + q_2 + q_3 = 1.92 + 21.0 + 2.63\text{ kJ} = 25.6\text{ kJ}.$

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14. Calculate the **standard enthalpy of formation**, ΔH_f^o, in kJ mol^{-1}, for **SO_2(g)** from the following data:

$S(s) + 3/2\ O_2(g) \longrightarrow SO_3(g)$ $\Delta H^o = -395.2$ kJ mol^{-1}
$2\ SO_2(g) + O_2(g) \longrightarrow 2\ SO_3(g)$ $\Delta H^o = -198.2$ kJ mol^{-1}

.

A) −10.6
B) −192.0
C) −89.1
D) −296.1
E) 192.0

The first reaction in the listed data is the formation reaction of SO_3(g). Therefore, we are given the enthalpy of formation of SO_3(g). The given enthalpy of the second listed reaction can also be written as

$$\Delta H^o = 2\ \Delta H_f^o[SO_3(g)] - 2\ \Delta H_f^o[SO_2(g)] \quad \text{(note that } \Delta H_f^o[O_2(g)] = 0)$$
$$= -198.2 \text{ kJ mol}^{-1} = 2 \times (-395.2 \text{ kJ mol}^{-1}) - 2\ \Delta H_f^o[SO_2(g)]$$

Therefore,

$$2\ \Delta H_f^o[SO_2(g)] = 2 \times (-395.2) - (-198.2) \text{ kJ mol}^{-1} = -592.2 \text{ kJ mol}^{-1}$$

or

$$\Delta H_f^o[SO_2(g)] = -592.2/2 \text{ kJ mol}^{-1} = -296.1 \text{ kJ mol}^{-1} .$$

15. PCl_5(s) can be prepared by the reaction $PCl_3(l) + Cl_2(g) \longrightarrow PCl_5(s)$. Calculate the **enthalpy change** (in **kJ**) that accompanies the production of **100.0 g** of PCl_5(s) by the above reaction, given the following data:

$P_4(s) + 6\ Cl_2(g) \longrightarrow 4\ PCl_3(l)$ $\Delta H^o = -1280$ kJ per mole of P_4(s)
$P_4(s) + 10\ Cl_2(g) \longrightarrow 4\ PCl_5(s)$ $\Delta H^o = -1774$ kJ per mole of P_4(s)

.

A) +134
B) −134
C) −78.1
D) +59.3
E) −59.3

The standard enthalpy of reaction for

$$PCl_3(l) + Cl_2(g) \longrightarrow PCl_5(s)$$

can be written as

$$\Delta H^o = \Delta H_f^o[PCl_5(s)] - \Delta H_f^o[PCl_3(l)] \quad \text{(note that } \Delta H_f^o[Cl_2(g)] = 0)$$

The enthalpies of formation in this formula equal ¼ the given enthalpies of reaction. Thus,

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$\Delta H^o = -1774/4 - (-1280/4)$ kJ mol^{-1} = -123.5 kJ mol^{-1} .

This is the enthalpy change per mole of $PCl_5(s)$ produced.
100.0 g of $PCl_5(s)$ corresponds to 100.0 g /(208.22 g mol^{-1}) = 0.4803 mol.

The enthalpy change associated with production of 0.9011 mol of $PCl_5(s)$ is

-123.5 kJ mol^{-1} × 0.4803 mol = -59.3 kJ .

16. A detector receives a signal consisting of green light, with a wavelength of 540 nm. The energy of the signal is 2.50 x 10^{-14} J. How many photons reach the detector?
.
A) 1.48 x 10^{7}
B) 6.79 x 10^{4}
C) 2.10 x 10^{-5}
D) 1.48 x 10^{4}
E) 3.25 x 10^{7}

$$E = \frac{hc}{\lambda} = \frac{(6.626x10^{-34}\,J.s)(3.00x10^{8}\,m.s^{-1})}{540x10^{-9}\,m} = 3.68x10^{-19}\,J \quad \text{is the energy of one photon}$$

The total energy is n times this where $n = \frac{2.50x10^{-14}\,J}{3.68x10^{-19}\,J} = 6.79x10^{4}$

17. Which of the following statements regarding quantum mechanics are **FALSE**?

(i) The energy of a photon is proportional to its frequency.
(ii) In a hydrogen atom, the electron is at a fixed distance from the nucleus.
(iii) As the velocity of a given particle gets larger, its wavelength gets shorter.
(iv) The size of an atomic orbital is mainly determined by the magnetic quantum number.
(v) For a given shell of a many-electron atom, d orbitals have higher energy than s orbitals.
.
A) i, iii
B) i, ii, v
C) iii, iv
D) i, iii, v
E) ii, iv

i) E = hν, where E is photon energy, ν is frequency and h is Planck's constant. The statement is TRUE.
ii) Electrons in the H atom have a probability distribution over a region of space which is spread over a range of distances from the nucleus. **The statement is FALSE.**

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iii) de Broglie's law states that momentum (and thus velocity, v) and wavelength (λ) are related by $mv = hc/\lambda$ so as v increases λ decreases. The statement is TRUE.
iv) m_ℓ, the magnetic quantum number does not determine size. Size is related to the n and ℓ quantum numbers. **The statement is FALSE.**
v) Within a shell (the n quantum number), energy is increases (it is less negative) as ℓ increases. The d orbitals (ℓ=2) are higher energy than s orbitals (ℓ=0). The statement is TRUE.

18. Which **one** of the following choices lists the species in order of increasing size?

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A) $I^- < I < Br$
B) $F^- < F < Cl$
C) $F^- < Cl^- < Cl$
D) $Cl^+ < Cl^- < Cl$
E) $F < F^- < Cl^-$

A) Anions are larger than neutral for the same element. $I^- > I$. The statement is FALSE
B) Anions are larger than neutral for the same element. $F^- > F$. The statement is FALSE.
C) Anions are larger than neutral for the same element. $Cl^- > Cl$. The statement is FALSE
D) Anions are larger than neutral for the same element. $Cl^- > Cl$. The statement is FALSE
E) Anions are larger than neutral for the same element and size increases down a group. **This statement is TRUE.**

19. The photoelectric effect is observed for a certain metal using light of wavelength 490 nm or smaller. What is the **minimum energy**, in J, that is required to eject **one** electron from the surface of this metal?

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A) 8.93×10^{-19}
B) 5.14×10^{-18}
C) 4.06×10^{-19}
D) 6.53×10^{-7}
E) 1.17×10^{3}

$$E = \frac{hc}{\lambda} = \frac{(6.626x10^{-34}\,J.s)(3.00x10^{8}\,m.s^{-1})}{490x10^{-9}\,m} = 4.06x10^{-19}\,J$$

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20. Which **one** of the following statements is **FALSE**?

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A) Nitrogen atoms in their ground state are paramagnetic.
B) $[He]2s^2$ is the electron configuration of the ground state of a Be atom.
C) Calcium atoms in their ground state are paramagnetic.
D) $[He]2s^2\ 2p^5$ is the electron configuration of the ground state of a F atom
E) $[Ar]4s^1 3d^1$ is the electron configuration of an excited state of a Ca atom.

A) N atom ground state electron configuration is $1s^2 2s^2 2p^3$ with the three 2p electrons unpaired. N is paramagnetic. The statement is TRUE.
B) Be atom ground state electron configuration is $1s^2 2s^2$. The statement is TRUE.
C) Ca atom ground state electron configuration is $[Ar]4s^2$. The two 4s electrons have opposite electron spin quantum numbers, and thus Ca is diamagnetic. **The statement is FALSE.**
D) F atom ground state electron configuration is $[He]2s^2\ 2p^5$. The statement is TRUE.
E) Ca atom ground state electron configuration is $[Ar]4s^2$. The 3d orbital is above 4s in Ca and thus $[Ar]4s^1 3d^1$ is an allowed configuration that is at higher energy than the ground state. The statement is TRUE.

21. Identify the **TRUE** statement(s):

(i) Be has a larger atomic radius than B.
(ii) Overall, electronegativity decreases from left to right across a period.
(iii) F^- has a smaller ionic radius than Na^+.
(iv) Rb has a lower first ionization energy than Na.
(v) All period 2 elements form acidic oxides.

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A) i, iv
B) i, ii, iv
C) ii, iii, v
D) i
E) iii, v

(i) Size increases from right to left across a period. Be is to the left of B, and thus Be has a larger atomic radius than B. **The statement is TRUE.**
(ii) Electronegativity increases from left to right across a period. The statement is FALSE.
(iii) The ground state configurations of F^- and Na^+ are both $[He]2s^2 2p^6$. The nuclear charge of Na is 2 larger than that of F. This will keep the electrons closer to the nucleus and Na^+ will be smaller than F^-. The statement is FALSE.
(iv) Ionization energies decrease going down a group. Rb is in period 5, while Na is in period 3. **The statement is TRUE.**
(v) The elements to the left of the periodic table (alkali and alkaline earth) form basic oxides. The statement is FALSE.

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22. C is the central atom in the thiocyanate anion, SCN^-. The **best** Lewis structure has the following **S, C and N** formal charges, respectively:

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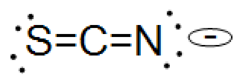
A) 0, +1, −2
B) −1, 0, 0
C) −2, +1, 0
D) −1, +1, −1
E) 0, 0, −1

Lewis structure of SCN^- is



The formal charges on S, C and N are 0, 0 and -1, respectively. Note that the minus formal charge goes to N rather than S (the structure with a single SC bond and a triple CN bond) because N is more electronegative than S.

23. How many **resonance structures** are required to portray the bonding in HSO_3^- ? (Sulfur is the central atom, is bonded only to oxygen atoms and has zero formal charge).

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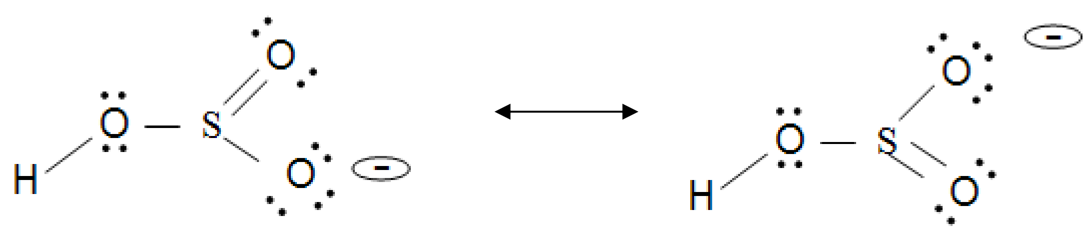
A) 6
B) 0
C) 1
D) 3
E) 2

Lewis structure of HSO_3^- is



TWO resonance structures are needed (not three, because the O-H oxygen is chemically different than the two O atoms only bonded to S).

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24. Based on Lewis structures with minimized formal charges, which of the following species would be expected to have the **longest** nitrogen-oxygen bond?

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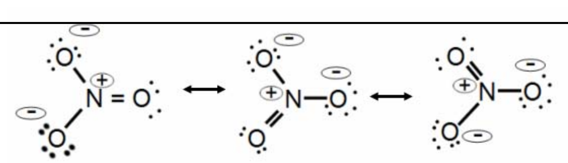
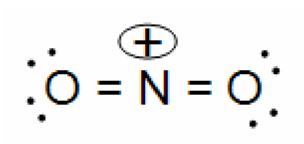
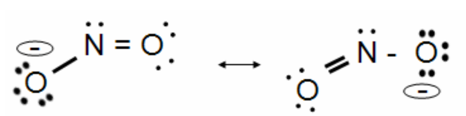
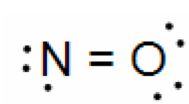
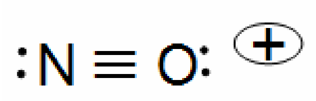
A) NO_3^-
B) NO_2^+
C) NO_2^-
D) NO
E) NO^+

The easiest way to answer this question is to realize that length of a bond is related to the ability of the two atoms to share their valence electrons – the more electrons that are being shared, the stronger and thus the shorter the bond is. From the N atom perspective, in this series the ability of the N to contribute to the bonding to oxygen decreases as the number of O atoms increase. From that point of view NO_3^- should have the longest N-O bond. In addition, adding one more -ve charge will increase electron – electron repulsion, further weakening the N-O bonds in NO_3^-.

The Lewis structures, with minimized formal charges are below. The nitrate ion has the lowest formal bond order, and thus the longest N-O bond length.

	STRUCTURE	bond length	bond order
A)		127 pm	1.33
B)		115 pm	2
C)		114 pm	1.5
D)		115 pm	2
E)		106 pm	3

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25. Choose the **one FALSE** statement about the Lewis structure of the peroxide anion, O_2^{2-}.
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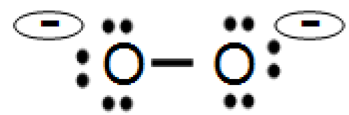
A) The oxygen-oxygen bond is a single bond.
B) Each oxygen obeys the octet rule.
C) Each oxygen has a formal charge of –1.
D) Two resonance forms are required to describe the bonding.
E) Each oxygen has 3 nonbonding electron pairs.

The Lewis structure of **peroxide anion** is



A) The statement is TRUE. The bond is single.
B) The statement is TRUE. There are 8 electrons in the valence shell of each O atom.
C) The statement is TRUE.
D) **The statement is FALSE. NO additional structures are required.**
E) The statement is TRUE.

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Extra space for rough work:

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- **Some general data are provided on this page.**
- **A Periodic Table with atomic weights is provided on the next page.**

R = 8.3145 J K^{-1} mol^{-1} = 0.08206 L atm K^{-1} mol^{-1}

$c = 2.9979 \times 10^{8}$ m s^{-1}

$m_e = 9.10 \times 10^{-31}$ kg

Specific heat of $H_2O(s)$ = 2.03 J / g·°C

Specific heat of $H_2O(l)$ = 4.18 J / g·°C

$N_A = 6.022 \times 10^{23}$ mol^{-1}

$h = 6.6256 \times 10^{-34}$ Js

density(H_2O, l) = 1.00g/mL

$\Delta H^{o}_{fus}[H_2O]$ = 6.01 kJ mol^{-1}

$\Delta H^{o}_{vap}[H_2O]$ = 44.0 kJ mol^{-1}

1 atm = 101.325 kPa = 760 mm Hg

1 J = 1 kg m^2 s^{-2} = 1 kPa L = 1 Pa m^3

1 cm^3 = 1 mL

1 Hz = 1 cycle/s

0°C = 273.15 K

1 m = 10^9 nm = 10^{10} Å

1 g = 10^3 mg

De Broglie wavelength:

$\lambda = h / mv = h / p$

Hydrogen atom energy levels:

$E_n = -R_H / n^2 = -2.178 \times 10^{-18}$ J / n^2

Solubility Guidelines for Common Ionic Solids

TABLE 5.1 Solubility Guidelines for Common Ionic Solids

Follow the lower-numbered guideline when two guidelines are in conflict. This leads to the correct prediction in most cases.

1. Salts of group 1 cations (with some exceptions for Li^+) and the NH_4^+ cation are soluble.
2. Nitrates, acetates, and perchlorates are soluble.
3. Salts of silver, lead, and mercury(I) are insoluble.
4. Chlorides, bromides, and iodides are soluble.
5. Carbonates, phosphates, sulfides, oxides, and hydroxides are insoluble (sulfides of group 2 cations and hydroxides of Ca^{2+}, Sr^{2+}, and Ba^{2+} are slightly soluble).
6. Sulfates are soluble except for those of calcium, strontium, and barium.

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