QUESTIONS

PART I MULTIPLE CHOICE

1. Which one of the following elements is an alkali metal?

A. Al   B. Cl   C. He
D. Na   E. O

2. Which of the following numbers has three significant figures?

A. 1,207   B. 4.250   C. 0.006
D. 0.0250   E. 0.03750

3. 1-butanethiol, one of the compounds giving skunks their distinctive odor, freezes at -115.7°C and boils at 98.5°C. What is its state of matter at 37°C, the normal body temperature of humans?

A. solid   B. liquid   C. gas
D. a mixture of solid and liquid   E. a mixture of liquid and gas

4. Which statement is **incorrect** according to modern atomic theory?

A. Some sugars have the formula $C_{3.5}H_7O_{3.5}$.
B. An atom of ^{14}N has the same approximate mass as an atom of ^{14}C.
C. Some atoms of ^{14}N behaves differently in chemical reactions than some atoms of ^{14}C.
D. An example of a chemical reaction is $CaCO_3 \rightarrow CaO + CO_2$.
E. C_2H_6 and C_3H_8 are both possible compounds of carbon and hydrogen.

5. The Law of Multiple Proportions states

A. Atoms of one element may combine in different ratios to form more than one compound.
B. Atoms of one element may combine in different ratios to form the same compound.
C. Atoms of two or more elements may combine in different ratios to form more than one compound.

D. Atoms of two or more elements may combine in different ratios to produce the same compound.
E. Compounds always form in the same percentages by mass.

6. Protons possess a ______ charge, and neutrons possess a ______ charge.

A. positive, zero
B. zero, positive
C. positive, negative
D. negative, positive
E. negative, negative

7. The atomic mass of an element is

A. The mass of the most abundant isotope of that element
B. The weighted average of the masses of the naturally occurring isotopes of that element
C. The arithmetic average of the masses of the isotopes of that element
D. The ratio of the mass of one atom of an isotope of that element to the mass of hydrogen
E. None of the above

8. In order to become a stable ion, how many electrons must an atom of nitrogen lose or gain?

A. lose 1
B. lose 2
C. lose 3
D. gain 5
E. gain 3

9. Xylene, a compound with the formula C_8H_{10}, is composed of

A. any combination of atoms of carbon and hydrogen in a four to five ratio
B. eight atoms of carbon and ten atoms of hydrogen
C. eight atoms of calcium and ten atoms of helium
D. ten atoms of carbon and eight atoms of hydrogen
E. any combination of atoms of carbon and hydrogen that add up to 18 total

10. Which one of the following reactions can be described as a double replacement reaction.

A. $Zn(s) + FeCl_2(aq) \rightarrow ZnCl_2(aq) + Fe(s)$
B. $CuSO_4(aq) + Fe(s) \rightarrow Cu(s) + FeSO_4(aq)$
C. $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
D. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
E. $CuCl_2(aq) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2AgCl(s)$

11. An atom containing 47 protons, 47 electrons and 60 neutrons has a mass number of

A. 13
B. 47
C. 60
D. 107
E. 154

12. Where is most of the mass of an atom concentrated?

A. Protons
B. Neutrons
C. Electrons
D. Nucleus

13. An imaginary element Xz consists of two isotopes having masses of 100.00 amu and 102.00 amu. A sample of Xz was found to consist of 75% ^{100}Xz and 25% ^{102}Xz. What is the atomic weight of Xz?

A. 100.25
B. 100.50
C. 101.75
D. 101.00
E. 101.50

14. Which pair of atoms constitutes a pair of isotopes of the same element?

A. $^{14}_{6}X$ and $^{14}_{7}X$
B. $^{14}_{6}X$ and $^{12}_{6}X$
C. $^{17}_{9}X$ and $^{17}_{8}X$
D. $^{19}_{10}X$ and $^{19}_{9}X$
E. $^{20}_{10}X$ and $^{20}_{11}X$

15. Which is the correct formula for the ionic compound containing Iron(III) ions and oxide ions?

A. FeO
B. FeO_2
C. Fe_2O_2
D. Fe_2O_3
E. Fe_3O_2

16. What is the name of $Mg_3(PO_4)_2$?

A. trimagnesium diphosphate
B. magnesium diphosphate
C. trimagnesium phosphate
D. magnesium phosphate
E. magnesium phosphorous oxide

17. One definition of an acid is a substance that provides which cation in water solution?

A. Na^+ B. H^+ C. OH^-
D. NH_4^+ E. Cl^-

18. The compound BrF_3 would be called

A. boron fluoride
B. triboron fluoride
C. tri (bromine fluoride)
D. boron trifluoride
E. bromine trifluoride

19. A chemical equation is balanced when

A. the total number of ions is the same in reactants and products.
B. the total number of molecules is the same in reactants and products.
C. the sum of the coefficients of the reactants is equal to the sum of the coefficients of the products.
D. the number of atoms of each element is the same in reactants and products.
E. the charge on each atom is the same in reactants and products.

20. The combination of ions most likely to produce a precipitate is

A. Li^+ and PO_4^{3-}
B. Pb^{2+} and NO_3^-
C. NH_4^+ and SO_4^{2-}
D. Fe^{3+} and OH^-
E. Mg^{2+} and $C_2H_3O_2^-$

21. The oxidation number of sulfur in calcium sulfate, $CaSO_4$, is

A. +6 B. +4 C. +2
D. 0 E. -2

22. When a substance loses electrons it is _____________; the substance itself is acting as a(n) ________________ agent.

A. oxidized, reducing
B. oxidized, oxidizing
C. reduced, oxidizing
D. reduced, reducing
E. dissolved, neutralizing

23. Which one of the following assumptions of the kinetic-molecular theory best explains the observation that a gas can be compressed?

A. Gas molecules move at random with no attractive forces between them.
B. The velocity of gas molecules is proportional to their Kelvin temperature.
C. The amount of space occupied by a gas is much greater than the space occupied by the actuual gas molecules.
D. In collisions with the walls of the container or with other molecules, energy is conserved.
E. Collisions with the walls of the container or with other molecules are elastic.

24. A mixture of N_2, O_2, and He gases has a total pressure of 760 mm Hg. If the partial pressures of N_2 is 90 mm Hg and of O_2 is 270 mm Hg. What is the partial pressure of He?

A. 1120 mm Hg B. 760 mm Hg C. 400 mm Hg
D. 360 mm Hg E. 120 mm Hg

25. The total pressure of a mixture of gases is

A. obtained by multiplying the individual pressures by the number of moles and averaging.
B. the sum of the partial pressures of the components.
C. dependent only upon the pressure of the gas which is present to the greatest extent.
D. the product of the individual pressures.
E. none of these.

26. A real gas will behave most like an ideal gas under conditions of __________.

A. high temperature and high pressure
B. high temperature and low pressure
C. low temperature and high pressure
D. low temperature and low pressure
E. STP

27. The change in the internal energy of a system that releases 2,500 J of heat and that does 7,655 J of work on the surroundings is __________ J.

A. -10,155 B. -5,155 C. -1.91×10^{7}

D. 10,155 E. 5,155

28. A __________ ΔH corresponds to an __________ process.

A. negative, endothermic
B. negative, exothermic
C. positive, exothermic
D. zero, exothermic
E. zero, endothermic

29. The quantity of heat needed to raise the temperature of 1 gram of a sample of a substance by 1°C is the sample's

A. heat capacity
B. specific heat
C. enthalpy
D. work
E. calorimetry

30. The ground state electron configuration for Zn is __________.

A. $[Kr]4s^2 3d^{10}$
B. $[Ar]4s^2 3d^{10}$
C. $[Ar]4s^1 3d^{10}$
D. $[Ar]3s^2 3d^{10}$
E. $[Kr]3s^2 3d^{10}$

31. Which one of the following represents an acceptable possible set of quantum numbers (in the order n, l, m_l, m_s) for an electron in an atom?

A. 2, 1, -1, 1/2
B. 2, 1, 0, 0
C. 2, 2, 0, 1/2
D 2, 0, 1, -1/2
E. 2, 0, 2, +1/2

32. There are __________ unpaired electrons in a ground state cobalt (Co) atom.

A. 0
B. 1
C. 2
D. 3
E. 4

33. An electron with n = 4 and l = 1 is classified as

A. a 4s electron
B. a 4d electron
C. a 4f electron
C. a 4p electron
D. none of these

34. The allowed m_l value(s) for a 2p electron is/are

A. 0
B. 0,1,2
C. 1,0,-1
D. 2,1,0,-1,-2
E. 3, 2, 1, 0, -1, -2, -3

35. Which one of the following is an incorrect subshell notation?

A. 4f B. 2p C. 3s
D. 1p E. 3d

36. In quantum mechanics, an orbital represents

A. the space occupied by an electron
B. the space occupied by the nucleus
C. the most probable location of an electron
D. the most probable location of the nucleus
E. the movement of an electron from one orbit to a different orbit

37. The ___________ subshell contains 5 orbitals.

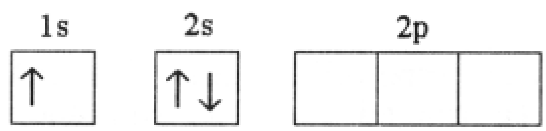
A. 4f B. 6f C. 4s
D. 3d E. 2p

38. The Pauli exclusion principle requires that

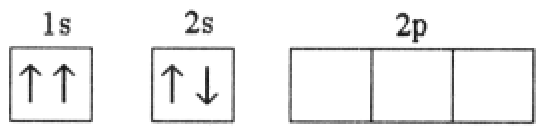
A. both the position of an electron and its momentum cannot simultaneously be known very accurately.
B. the wavelength and mass of a subatomic particle are related by $\lambda = h/mv$.
C. the wavelength of a photon of light times its frequency is equal to the speed of light.
D. no two electrons in the same atom can have the same set of four quantum numbers.
E. an electron can have either particle character or wave character.

39. Which electron configuration represents a violation of the Pauli exclusion principle?

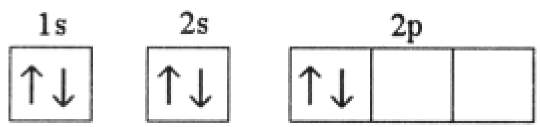
A.



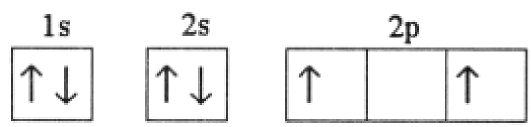
B.



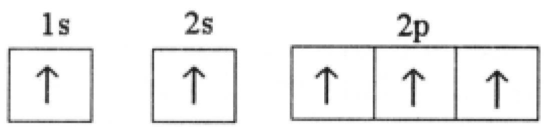
C.



D.



E.



40. From a consideration of electronic configurations, which of the elements indicated below would be found in Group 7A?

A. $1s^2, 2s^2, 2p^2$
B. $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$
C. $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$
D. $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^5, 4s^2$
E. $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6$

41. Atomic radius generally increases as we move __________.

A. down a group and from right to left across a period
B. up a group and from left to right across a period
C. down a group and from left to right across a period
D. up a group and from right to left across a period
E. down a group; the period position has no effect

42. Of the following elements, which has the smallest <u>first</u> ionization energy?

A. Na B. Al C. Mg
D. Cl E. Si

43. The group in the peridic table which has a total valence electron configuration of ns^2 np^3 is

A. the group 2a elements
B. the group 3A elements
C. the group 4a elements
D. the group 5A elements
E. the group 6a elements

44. In liquids, the attractive intermolecular forces are __________.

A. very weak compared with kinetic energies of the molecules
B. strong enough to hold molecules relatively close together
C. strong enough to keep the molecules confined to vibrating about their fixed lattice points
D. not strong enough to keep molecules from moving past each other
E. strong enough to hold molecules relatively close together but not strong enough to keep molecules from moving past each other

45. Which one of the following derivatives of ethane has the highest boiling point?

A. C_2Br_6
B. C_2F_6
C. C_2I_6
D. C_2Cl_6
E. C_2H_6

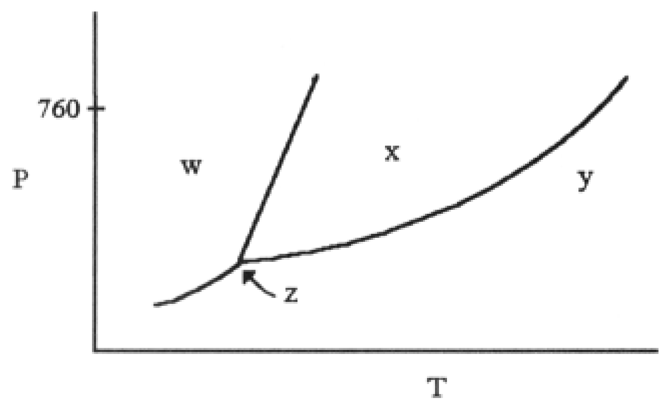
46. The predominant intermolecular force in $(CH_3)_2NH$ is __________.

A. London dispersion forces
B. ion-dipole forces
C. ionic bonding
D. dipole-dipole forces
E. hydrogen bonding

47. The ease with which the charge distribution in a molecule can be distorted by an external electrical field is called the __________.

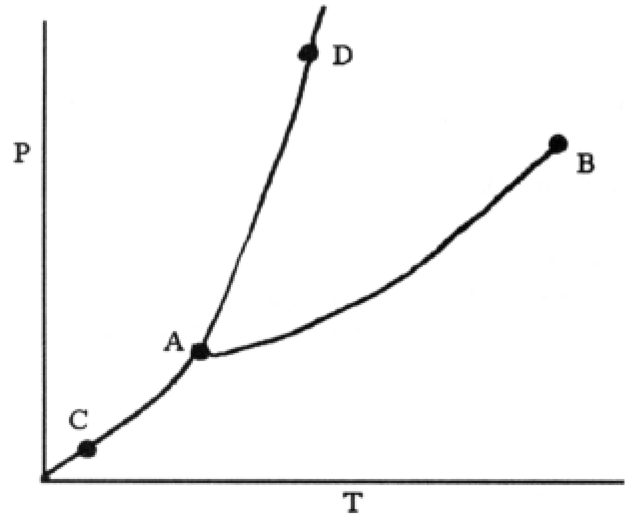
A. electronegativity
B. hydrogen bonding
C. polarizability
D. volatility
E. viscosity

48. The phase diagram of a substance is given above. The region that corresponds to the liquid phase is __________.



A. w
B. x
C. y
D. z
E. x and y

49. On the phase diagram shown, the coordinates of point __________ correspond to the critical temperature and pressure.



A. A B. B C. C
D. D E. E

50. Which of the following species is isoelectronic with Kr?

A. Xe B. K^+ C. In^{3+}
D. S^{2-} E. Sr^{2+}

51. An electron with n = 6 and l = 0 is classified as a

A. 6s electron B. a 6d electron C. a 6g electron
C. 6p electron D. a 6f electron

52. From a consideration of electronic configurations, which of the elements indicated below would be found in Group 8A?

A. $1s^2, 2s^2, 2p^2$
B. $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$
C. $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$
D. $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^5, 4s^2$
E. $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6$

53. How many electrons can occupy the 3d subshell?

A. 6 B. 2 C. 10

D. 4 E. 8

54. The larger the difference in electronegativity

1. the more ionic the bond.
2. the more covalent the bond.
3. the more polar the bond.

a. 1 only b. 2 only c. 3 only
d. 1 and 3 only e. 2 and 3 only

55. The Lewis structure for ammonia, NH_3, has

a. four bonding pairs.
b. two bonding pairs and two lone pairs.
c. three bonding pairs and one lone pair.
d. one bonding pair and three lone pairs.
e. four lone pairs.

56. Which of the following statements is not true about resonance?

a. A single Lewis formula does not provide an adequate representation of the bonding.
b. Resonance describes the oscillation and vibration of electrons.
c. Resonance describes a more stable situation than does any one contributing resonance formula.
d. Resonance describes the bonding as intermediate between the contributing resonance formulas.
e. The contributing resonance formulas differ only in the arrangement of the electrons.

57. How many pairs of electrons are there around the central atom in the Lewis structure of SCl_4?

A. 2 B. 3 C. 4
D. 5 E. 6

58. Which one of the following statements best describes the Lewis structure of NO_2^-?

A. There are 2 single N – O bonds and no lone pairs on the nitrogen atom.

B. There are 2 single N – O bonds and one lone pair on the nitrogen atom.
C. There are 2 single N – O bonds and two lone pairs on the nitrogen atom.
D. There is 1 single N – O bond, 1 double N – O bond and one lone pair on the nitrogen atom
E. There are 2 double N – O bonds and no lone pairs on the nitrogen atom.

59. Atoms having an electronegativity difference greater than 2 units are expected to form

A. No bonds
B. polar covalent bonds
C. non-polar covalent bonds
D. ionic bonds
E. covalent bonds

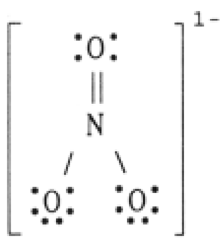
60. What is the maximum number of double bonds that a hydrogen atom can form?

A. 0
B. 1
C. 2
D. 3
E. 4

61. A double bond consists of __________ pairs of electrons shared between two atoms.

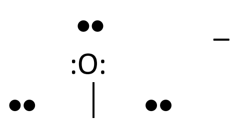
A. 1
B. 2
C. 3
D. 4
E. 6

62. The formal charge on nitrogen in NO_3^- is __________.



A. -1
B. 0
C. +1
D. +2
E. -2

63. The Lewis structure for the BrO_3^- ion is shown below:



:O — Br — O:

•• •• ••

The molecular geometry for this polyatomic ion is

A. linear B. tetrahedral C. bent
D. trigonal planar E. pyramid

64. The Lewis structure of BrO_3^- is given in the previous question. The electronic geometry of BrO_3^- is _______________ and the overall molecule is ____________.

A. tetrahedral, polar B. tetrahedral, nonpolar
C. trigonal planar, polar D. bent, polar
E. pyramidal, polar

65. In the ICl_4^- ion, how many lone pairs of electrons are there about the central iodine atom?

a. 0 b. 1 c. 2
d. 3 e. 4

66. All of the following molecules have polar bonds and are polar molecules except

a. ICl_3 b. BCl_3 c. PCl_3
d. H_2CCl_2 e. ICl.

67. The hybridization of the central atom in a molecule is described as sp^3. The arrangement in space of the hybrid orbitals about that atom is

a. linear. b. trigonal planar. c. tetrahedral.
d. trigonal bipyramidal. e. octahedral.

68. The H-N-H angle in ammonia

a. 180^o b. 120^o c. slightly less than 120^o
d. 109^o e. slightly less than 109^o

69. How many σ (sigma bonds) are there in a triple bond?

A. 1 B. 2 C. 3
D. 4 E. 5

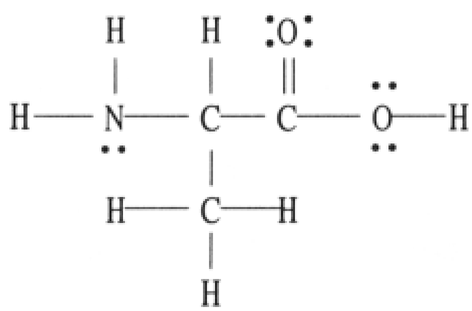
70. The idealized bond angle in the trigonal planar electronic geometry is

A. 90^o
B. 120^o
C. 109.5^o
D. 180^o
E. None of these angles

71. According to VSEPR theory, if there are five electron domains in the valence shell of an atom, they will be arranged in a(n) __________ geometry.

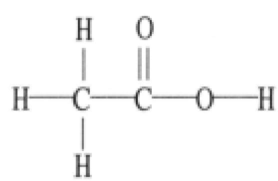
A. octahedral
B. linear
C. tetrahedral
D. trigonal planar
E. trigonal bipyramidal

72. There is/are __________ σ bond(s) in the molecule below.



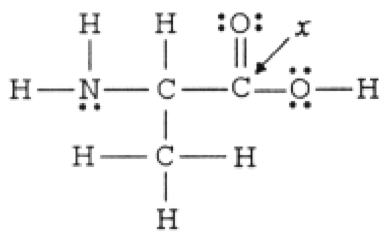
A. 1
B. 2
C. 12
D. 13
E. 18

73. The molecular geometry of the left-most carbon atom in the molecule below is __________.



A. trigonal planar
B. trigonal bipyramidal
C. tetrahedral
D. octahedral
E. T-shaped

74. The hybridization of the carbon atom labeled x in the molecule below is __________.



A. sp B. sp^2 C. sp^3

D sp^3d E. sp^3d^2

PART II

1. Balance the following equations:

A. __$CuSO_4$ + __Al → __$Al_2(SO_4)_3$ + __Cu

B. __$C_8H_{18}(g)$ + __$O_2(g)$ → __$CO_2(g)$ + __$H_2O(l)$

2. Complete and balance the following reactions.

A. __$Pb(NO_3)_2$ + __NaBr →

B. __H_3PO_4 + __NaOH →

3. Write balanced net ionic equations for the following balanced reactions. If no reaction occurs, write NR.

A. $Ca(NO_3)_2(aq)$ + $K_2CO_3(aq)$ → $CaCO_3(s)$ + 2 $KNO_3(aq)$

B. $Na_2CO_3(aq)$ + 2 HCl(aq) → 2 NaCl(aq) + $H_2O(l)$ + $CO_2(g)$

PART III PROBLEMS - each as indicated

FOR FULL CREDIT ON THE PROBLEMS, SHOW YOUR WORK IN AN ORGANIZED MANNER.

1. A small bottle contains 2.17 mL of a red liquid. The total mass of the bottle and the liquid is 5.261 g. The empty bottle weighs 3.006 g. What is the density of the liquid (in g/mL)

2. The daily dose of ampicillin for the treatment of an ear infection is 115 mg of ampicillin per kilogram of body weight (115 mg/kg of body weight). What is the daily dose for a 34 pound child?

3. Ethyl mercaptan is an odorous substance added to natural gas to make leaks easily detectable. Analysis of a sample of this substance determined that this compound contains by weight:

38.65% carbon 9.75% hydrogen 51.6% sulfur (S)

Determine the empirical formula of ethyl mercaptan

4. What volume (in mL) of 0.300 M $CaCl_2$ is needed to react completely 40.0 mL of 0.2 M H_3PO_4 solution? The balanced equation is

$$3\ CaCl_2(aq) + 2\ H_3PO_4(aq) \rightarrow Ca_3(PO_4)_2\ (aq) + 6\ HCl(l)$$

5. A sample of an ideal gas has a volume of 1.50 L at 23°C. What will be its volume (in liters) if the temperature is increased to 223°C at constant pressure?

6. Darvon ($C_{22}H_{30}ClNO_2$) in the past has been used as a prescription painkiller.

A. What is the molar mass of darvon?

B. What is percent by weight of carbon in darvon?

C. How many moles of darvon are present in a 100.0 g sample.

D. How many molecules of darvon are present in a 100.0 g sample?

E. How many hydrogen atoms are present in a 100.0 g sample of darvon?

7. Consider the following balanced equation:

$$6\ I_2O_5 + 20\ BrF_3 \rightarrow 12\ IF_5 + 15\ O_2 + 10\ Br_2$$

Molar masses: I_2O_5 = 334 g/mole, BrF_3 = 137 g/mole, IF_5 = 222 g/mole, Br_2 = 160 g/mole, O_2 = 32.0 g/mole

If 500.0 g of I_2O_5 and 500.0 g of BrF_3 powder are reacted:

A. How many grams of IF_5 can be produced?

B. Which substance I_2O_5 or BrF_3 is the limiting reactant? I_2O_5 or BrF_3

C. Which substance is present in excess? I_2O_5 or BrF_3

D. How many grams of the substance left remains unreacted?

E. If only 297.0 g of IF_5 actually produced, what is the percent yield of IF_5?

8. Ammonium nitrate, NH_4NO_3, can decompose explosively when heated to a high temperature according to the following balanced equation:

$$2\ NH_4NO_3(s) \rightarrow 2\ N_2(g) + 4\ H_2O(g) + O_2(g)$$

If a 100.0 g sample of NH_4NO_3 decomposes at 450°C, how many liters of gaseos products will be formed? Assume the atmospheric pressure is 1 atm and the temperature of the gases formed is 450°C.

9. Calculate the density of SCl_2 gas (in g/L) at 37°C and 5.00 atm pressure.

10. Draw Lewis structure for the following species.

A. SO_2

B. IO_3^-

C. BrF_4^-

D. HCN

11. Circle any substance listed below that exhibits resonance.

SO_2 IO_3^- BrF_4^- HCN

12. Circle any substance that is polar.

SO_2 IO_3^- BrF_4^- HCN

13. Complete the following table (using your Lewis structures from question 7)

SUBSTANCE	ELECTRONIC GEOMETRY	MOLECULAR GEOMETRY	HYBRIDIZATION
SO_2			
IO_3^-			
BrF_4^-			
HCN			

14. What is the wavelength (in meters) of a photon that has an energy of 4.38×10^{-18} J?

15. (16 Points) The specific heat of liquid bromine is 0.226 J/g °C. How much heat (J) is required to raise the temperature of 10.0 mL of bromine from 25.00 °C to 27.30 °C? The density of liquid bromine: 3.12 g/mL.

16. (16 Points) Calculate the standard heat of formation, ΔH_f of $C_2H_5OH(l)$ from the measured heat of combustion of $C_2H_5OH(l)$ and the other tabulated ΔH_f values given below:
ΔH_f CO_2 (g) = -394 kJ/mol and ΔH_f $H_2O(l)$ = -286 kJ/mole

$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$ ΔH = -1368 kJ

17. (16 Points) Given the following reactions

$Fe_2O_3(s) + 3\ CO(g) \rightarrow 2\ Fe(s) + 3\ CO_2(g)$ ΔH = -28.0 kJ
$3\ Fe(s) + 4\ CO_2(g) \rightarrow 4\ CO(g) + Fe_3O_4(s)$ ΔH = +12.5 kJ

the enthalpy of the reaction of Fe_2O_3 with CO

$$3Fe_2O_3(s) + CO(g) \rightarrow CO_2(g) + 2Fe_3O_4(s)$$

is ___________ kJ.

18. (16 Points) From a consideration of the reaction

$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$ ΔH = -192.8 kJ

if 2.00×10^2 g of SO_3 were produced, then the amount of heat released should be

SOME USEFUL CONVERSION FACTORS

1 ft = 12 in	1 in = 2.54 cm
1 liter = 1.0577 quart	1 gallon = 4 quart
1 mile = 1.61 kilometers	1 meter = 100 centimeters
1 liter = 1000 mL	1 cm^3 = 1 ml
1 barrel = 42.0 gallons	1 quart = 9.46×10^{-4} m^3

1 g = 1000 mg
1 quart = 32 fluid ounces
1 torr = 1 mm Hg
1 kg = 2.2 lbs

1 kg = 1000 g
1 g = 1000 mg
1 atm = 760 mm Hg = 760 torr

FORMULAS AND CONSTANTS

Avogadro's Number = 6.02×10^{23} anythings/mole

STP: 1 atm and $0^{o}C$

D = M/V

$$\text{gram} \xrightarrow{\div \text{ MW or AW}} \text{mole} \xrightarrow{\times\, 6.02 \times 10^{23}} \text{molecule} \xrightarrow{\times \text{ subscript}} \text{atoms}$$

$$\text{atoms} \xrightarrow{\div \text{ subscript}} \text{molecules} \xrightarrow{\div\, 6.02 \times 10^{23}} \text{mole} \xrightarrow{\times \text{ MW or AW}} \text{grams}$$

$$\text{Grams A} \xrightarrow{\div \text{ AW or MW}} \text{mole A} \xrightarrow[\text{balanced equation}]{\text{coefficients of}} \text{mole B} \xrightarrow{\times \text{ AW or MW}} \text{grams B}$$

$$R = 0.0821 \frac{L \bullet atm}{mole \bullet K}$$

PV = nRT

$$M = \frac{DRT}{P} = \frac{gRT}{PV}$$

$$\frac{P_2V_2}{P_1V_1} = \frac{n_2RT_2}{n_1RT_1}$$

$^{o}C = K - 273$

$^{o}K = {}^{o}C + 273$

$^{o}F = 9/5(^{o}C) + 32$

$^{o}C = 5/9(^{o}F - 32)$

SOLUBILITY RULES

NO_3^- All nitrates are soluble

$C_2H_3O_2^-$ All acetates are soluble

ClO_3^- All chlorates are soluble

Cl^- All chlorides are soluble except $AgCl$, Hg_2Cl_2 and $PbCl_2$

Br^- All bromides are soluble except $AgBr$, Hg_2Br_2, $PbBr_2$ and $HgBr_2$

I^- All Iodides are soluble except AgI, Hg_2I_2, PbI_2 and HgI

SO_4^{2-} All sulfates are soluble except $CaSO_4$, $SrSO_4$, $BaSO_4$, $PbSO_4$, Hg_2SO_4 and Ag_2SO_4

S^{2-} All sulfides are insoluble except those of the group IA and group IIA elements and $(NH_4)_2S$

CO_3^{2-} All carbonates are insoluble except those of the group IA elements and $(NH_4)_2CO_3$

SO_3^{2-} All sulfites are insoluble except those of the group IA elements and $(NH_4)_2SO_3$

PO_4^{3-} All phosphates are insoluble except those of the group IA elements and $(NH_4)_3PO_4$

OH^- All hydroxides are insoluble except those of the group IA elements and $Ba(OH)_2$, $Sr(OH)_2$ and $Ca(OH)_2$

USEFUL INFORMATION

1s

2s 2p

3s 3p 3d

4s 4p 4d 4f

5s	5p	5d	5f
6s	6p	6d	6f
7s	7p	7d	7f

$\Delta E = q + w$

USEFUL INFORMATION

$1 \text{ nm} = 1 \times 10^{-9} \text{ m}$

$1 \text{ liter} = 1000 \text{ mL}$

$c = 3.0 \times 10^{8} \text{ m/sec}$

$1 \text{ mL} = 1 \text{ cm}^3$

$h = 6.626 \times 10^{-34} \text{ joule sec}$

Avogadro's Number = 6.02×10^{23}

$B = 2.179 \times 10^{-18} \text{ joule}$

$\Delta E = q + w$

$c = \lambda \nu$

$\Delta E = h\nu$

q = (number of grams)(specific heat)(ΔT)

$$\Delta H^{o}_{rx} = \sum \Delta H^{o}_{f}(products) - \sum \Delta H^{o}_{f}\ (\text{reactants})$$

$$\Delta H^{o}_{rx} = \sum \text{Energies of bonds broken} - \sum \text{Energies of bonds formed}$$

$$\Delta E = B\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

Solution

ANSWERS to CHM1011 PRACTICE FINAL EXAM FALL 2016

Multiple Choice Questions

1.	D	10.	E	19.	D
2.	D	11.	D	20.	D
3.	B	12.	D	21.	A
4.	A	13.	B	22.	A
5.	C	14.	B	23.	C
6.	A	15.	D	24.	C
7.	B	16.	D	25.	B
8.	E	17.	B	26.	B
9.	B	18.	E		

27. A — q = -2500 J   w = -7655
$\Delta E = q + w = (-2500) + (-7655) = -10155$ J

28. B

29. B

30. B — Zn has 30 electrons so $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$ or [Ar] $4s^2\ 3d^{10}$

31. A — n has to be a positive integer.
n has to be greater than *l*. *l* must be an integer between 0 and n-1.
m_l has to be +*l* and –*l*
mm_s can only be + ½ or – ½
Answer B is wrong because $m_s = 0$
Answer C is wrong because *l* can not equal n.
Answers D and E are wrong if *l* = 0 than m_l can only equal 0.

32. D — Cobalt has 27 electrons. Its' electron configuration is
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^7$
The only partially filled orbital is the 3d subshell.
It has two orbitals doubly filled and three orbitals singly filled, so
Three unpaired electrons.

33. C — n refers to the principal quantum number in this case an orbital in the 4^{th} shell. The *l* quantum number denotes the subshell.

If $l = 0$ a s subshell or orbital $l = 1$ a p subshell or orbital
$l = 2$ a d subshell or orbital $l = 3$ a f subshell or orbital
In this question $l = 1$ so this refers to a 4p electron

34. C m_l has to being within the regions of $+l$ to $-l$. For a 2p electron l is Equal to 1 so m_l can only be 1, 0 or -1.
35. D The n =1 shell has only 1 subshell the 1s.
36. C
37. D d subshells contain 5 orbitals (s has 1 orbital, p has 3 orbitals and f have 7 orbitals)
38. D
39. B Whenever two electrons present in the same orbital, they must have opposite spins (represented by arrows either up or down).
40. B The group 7A elements have a valence electron configuration of ns^2 np^5. Answer b is the only answer which has this valence electron configuration.
41. A As one goes across the periodic table screening of electrons causes atoms to get slightly smaller. As you down a column electrons are added to a shell farther away from the nucleus so atom gets bigger.
42. A Ionization increases across rows and decreases as you go down columns. All the elements listed here are in the third row of the periodic table so Na is the smallest and Cl should be the largest.
43. D Simply count up the superscripts to get the group number. So ns^2 np^3 is 2 + 3 = 5 so group 5A.
44. E In liquids molecules are able to move around relative to one another but not break free from each other.
45. C The molecule with the largest molar mass will have the strongest intermolecular forces and thus the highest boiling point.
(C_2Br_6 – 504 g/mole, C_2F_6 – 138 g/mole, C_2I_6 – 786 g/mole, C_2Cl_6 – 237 g/mole, C_2H_6 – 30 g/mole).
46. E This molecule has hydrogen bonds, dipole-dipole forces and London dispersion forces. Hydrogen bonds are a more important intermolecular force than the other 2.
47. C This is the definition of polarizability
48. B Region x corresponds to the liquid region
49. B The critical temperature and pressure correspond to the temperature and pressure where the liquid-vapor curve ends. In this drawing it is point B.

44. E Nonpolar bonds form when two atoms of equal electronegativity bond. Typically this means the atoms must be the same.

45. E Phosphorous (P) is in group 5A. It needs to gain 3 electrons to reach a noble gas number of electrons. If it gains 3 electrons it will form a P^{3+} ion.

46. C This is the definition of electronegativity.

47. B For oxygen there are 6 valence electrons.
The central oxygen has a single lone pair it gets 2 electrons from it.
It also is involved in three bonds (1 double bond and 1 single bond). It gets 1 electron from each bond for a total of three.
Oxygen owns 5 electrons in this structure. The formal charge is thus 6 – 5 = +1.

48. C Covalent bonds typically form when nonmetallic atoms form bonds.

49. B Resonance requires that multiple Lewis structures must be drawn and the this indicates that the actual structure is an average of the resonance forms.

50. E Kr has 36 electrons. Only Sr^{2+} has 36 electrons.

51. C

52 E ns^2 np^6 configuration corresponds to 8A (or 18)

53. C 5 d orbitals so it can hold 10 electrons

54. D The larger the difference in electronegativity the more polar the bond. The more polar the bond, the more ionic it is.

55. C The Lewis structure for ammonia has 3 bonds and a lone pair.

56. B Resonance structures requires multiple Lewis structure be drawn. The actual structure is an average of all the Lewis structures.

57. D SCl_4 has 4 bonding regions and a lone pair.

58. D The actual structure averages the single and double bond but cannot be drawn.

59. D When the electronegativity difference is greater than 2, the bond is considered ionic.

60. A Hydrogen only wants 2 electrons to get to its nearest noble gas so it cannot form anything other than a single bond.

61. B In a double bond four electrons are shared (2 electron pairs).

62. C Formal charge on N is 5 = 0 – 1/2 (8) = +1

63. E Br has 3 bonding regions and one lone pair. It has tetrahedral electron domain geometry and pyramidal molecular geometry.

64.	A	It is polar because it has a lone pair on Br.
65.	C	There are 12 electrons around the I, four bonding pairs and 2 lone Pairs
66.	B	BCl_3 It is the only structure that does not have a lone pair on the central atom or unsymmetrical bonding.
67.	C	Sp^3 hybridization always has tetrahedral electron domain geometry.
68.	E	The lone pair on NH_3 reduces the typical tetrahedral bond angle slightly.
69.	A	Every bond consists of only one sigma bond.
70.	B	
71.	E	Five electron domains corresponds to trigonal bipyramid electron domain geometry.
72.	C	13 atoms so 12 sigma bonds
73.	C	4 bonds and no lone pairs, so tetrahedral electron domain and molecular geometry.
74.	B	sp^2 hybridizations because there are three electron domains.

Part II

1. A. $3\ CuSO_4 + 2\ Al \rightarrow Al_2(SO_4)_3 + 3\ Cu$

   B. $2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$

2. A. $Pb(NO_3)_2 + 2\ NaBr \rightarrow PbBr_2 + 2\ NaNO_3$

   B. $H_3PO_4 + 3\ NaOH \rightarrow Na_3PO_4 + 3H_2O$

3. A. $Ca(NO_3)_2\,(aq) + K_2CO_3\,(aq) \rightarrow CaCO_3\,(s) + 2\ KNO_3\,(aq)$

   Soluble Soluble Insoluble Soluble

$Ca^{2+}(aq) + 2NO_3^-(aq) + 2K^+(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s) + 2K^+(aq) + 2NO_3^-(aq)$

Spectator ions are $2K^+(aq) + 2NO_3^-(aq)$

Net ionic equation is: $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

B. $Na_2CO_3(aq)$ + $2HCl(aq)$ → $2\ NaCl(aq)$ + $H_2O(l)$ + $CO_2(g)$

Soluble strong Acid Soluble Non Non

$2\ Na^+(aq) + CO_3^{2-}(aq) + 2\ H^+(aq) + 2\ Cl^-(aq) \rightarrow 2\ Na^+(aq) + 2\ Cl^-(aq) + H_2O(l) + CO_2(g)$

Spectator ions are $Na^+(aq)$ and $Cl^-(aq)$

Net ionic equation is: $2\ H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$

Part III

1. Mass of liquid = 5.261 g – 3.006 g = 2.255 g

Volume of liquid = 2.17 mL

$$D = \frac{2.255\ \text{g}}{2.17\ \text{mL}} = 1.04\ \text{g/mL}$$

2. $$34\ \text{lbs}\ \frac{1\ \text{kg}}{2.2\ \text{lbs}}\ \text{x}\ \frac{115\ \text{mg ampicillin}}{1\ \text{kg body weight}} = 1777\ \text{mg}$$

To 2 significant figures 1800 mg or 1.8 g

3. $$\text{mole carbon} = 38.65\ \text{g C x}\ \frac{1\ \text{mole C}}{12\ \text{g C}} = 3.22\ \text{mole C}$$

$$\text{mole hydrogen} = 9.75\ \text{g H x}\ \frac{1\ \text{mole H}}{1.01\ \text{g H}} = 9.65\ \text{mole H}$$

$$\text{mole sulfur} = 51.6\ \text{g S x}\ \frac{1\ \text{mole S}}{32.06\ \text{g S}} = 1.61\ \text{mole S}$$

Divide each mole by the smallest.

$C_{\frac{3.22}{1.61}}H_{\frac{9.65}{1.61}}S_{\frac{1.61}{1.61}}$ → C_2H_6S (this is the empirical formula of ethyl mercaptan.

4. $3\ CaCl_2 + 2\ H_3PO_4 \rightarrow Ca_3(PO_4)_2 + 6\ HCl$
0.3 M 0.2 M
? mL 40 mL

Step 1: Find mole H_3PO_4
Step 2 Convert mole H_3PO_4 to mole $CaCl_2$
Step 3 Calculate volume $CaCl_2$

$$\text{mole } H_3PO_4 = M \text{ x } L = \frac{0.2 \text{ mole}}{L} \text{ x } 0.040 \text{ L} = 0.008 \text{ mole } H_3PO_4$$

$$\text{mole } CaCl_2 = 0.008 \text{ mole } H_3PO_4 \text{ x } \frac{3 \text{ mole } CaCl_2}{2 \text{ mole } H_3PO_4} = 0.012 \text{ mole } CaCl_2$$

Volume $CaCl_2$ needed (L) = $\frac{\text{mole}}{M} = \frac{0.012 \text{ mole}}{0.300 \text{ mole/L}} = 0.04 \text{ L or } 40 \text{ mL}$

5. $V_1 = 1.5$ L $T_1 = 23^oC = 296$ K

$V_2 = ?$ $T_2 = 223^oC = 496$ K

Moles (n) and pressure (P) are constant

$$\frac{P_2V_2}{P_1V_1} = \frac{n_2RT_2}{n_1RT_1}$$ so P_2 and P_1, n_2 and n_1 and R all cancel

Leaving $$\frac{V_2}{V_1} = \frac{T_2}{T_1}$$

Plugging in we get

$$\frac{V_2}{1.5 \text{ L}} = \frac{496 \text{ K}}{296 \text{ K}}$$ Solving for V_2

$$V_2 = \frac{(496 \text{ K})(1.5 \text{ L})}{296 \text{ K}} = 2.51 \text{ L}$$

6. A. 22 C x 12.0 = 264
30 H x 1.0 = 30

1 Cl x 35.5 = 35.5
1 N x 14.0 = 14.0
2 O x 16.0 = <u>32.0</u>
375.5 g.mole

B. $\% \text{ carbon} = \frac{264}{375.5} \text{ x } 100 = 70.3\% \text{ C}$

C. $\text{mole darvon} = 100.0 \text{ g darvon x } \frac{1 \text{ mole darvon}}{375.5 \text{ g darvon}} = 0.266 \text{ mole darvon}$

D. $\text{molecules darvon} = 0.266 \text{ mole darvon x } \frac{6.02 \text{ x } 10^{23} \text{ molecules}}{1 \text{ mole darvon}} =$

1.60 x 10^{23} molecules darvon

E. $\text{hydrogen atoms} = 1.60 \text{ x } 10^{23} \text{ molecules darvon x } \frac{30 \text{ atoms H}}{1 \text{ molecule darvon}} =$

4.80 x 10^{24} hydrogen atoms

7. A. Maximum amount of IF_5 produced by reaction of 500.0 g I_2O_5

$$500.0 \text{ g } I_2O_5 \text{ x } \frac{1 \text{ mole } I_2O_5}{334 \text{ g } I_2O_5} \text{ x } \frac{12 \text{ mole } IF_5}{6 \text{ mole } I_2O_5} \text{ x } \frac{222 \text{ g } IF_5}{1 \text{ mole } IF_5} = 664.7 \text{ g } IF_5$$

Maximum amount of IF_5 produced by reaction of 500.0 g BrF_3

$$500.0 \text{ g } BrF_3 \text{ x } \frac{1 \text{ mole } BrF_3}{137 \text{ g } BrF_3} \text{ x } \frac{12 \text{ mole } IF_5}{20 \text{ mole } BrF_3} \text{ x } \frac{222 \text{ g } IF_5}{1 \text{ mole } IF_5} = 486.1 \text{ g } IF_5$$

486.1 < 664.7 so 486.1 g of IF_5 is the maximum amount of IF_5 that can be made.

B. BrF_3 is the limiting reactant (it produces the smaller quantity of IF_5)

C. I_2O_5 is in excess.

D. First find g I_2O_5 needed and subtract this amount from 500.0 g.

$$500.0 \text{ g BrF}_3 \text{ x } \frac{1 \text{ mole BrF}_3}{137 \text{ g BrF}_3} \text{ x } \frac{6 \text{ mole I}_2\text{O}_5}{20 \text{ mole BrF}_3} \text{ x } \frac{334 \text{ g I}_2\text{O}_5}{1 \text{ mole I}_2\text{O}_5} = 365.7 \text{ g I}_2\text{O}_5 \text{ needed}$$

Mass of I_2O_5 unreacted is 500.0 – 365.7 = 134.3 g I_2O_5 left unreacted.

8. In this problem, the data indicates that we will need to use the ideal gas law. However, we can only use mole of gases when we use the ideal gas law.
   The first thing we need to do is to convert the 100.0 g of NH_4NO_3 into mole of gases.

   To do this

   First: find the mole of NH_4NO_3 present.
   Second: Convert mole NH_4NO_3 to mole of total gases present.
   From the balanced equation we see that 2 mole NH_4NO_3 produces 2 mole of N_2gas, 4 mole of H_2O gas and 1 mole of O_2 gas.

   That is 2 mole NH_4NO_3 produces 7 moles of gases.

   mole of ammonium nitrate = 100 g NH_4NO_3 x $\frac{1 \text{ mole NH}_4\text{NO}_3}{80 \text{ g NH}_4\text{NO}_3}$ = 1.25 mole

$$\text{mole of gas formed} = 1.25 \text{ mole NH}_4\text{NO}_3 \text{ x } \frac{7 \text{ mole of gases}}{2 \text{ mole NH}_4\text{NO}_3} = 4.375 \text{ mole gas}$$

   We can now use this mole of gas to plug into the ideal gas law to determine the total volume occupied by the gases.

   n = 4.375 mole T = 450^{o}C =723 K P = 1 atm

   R = $0.0821 \frac{\text{L} \bullet \text{atm}}{\text{mole} \bullet \text{K}}$

$$V = \frac{nRT}{P} = \frac{(4.375 \text{ mole})(0.0821 \frac{\text{L} \bullet \text{atm}}{\text{mole} \bullet \text{K}})(723 \text{ K})}{1 \text{ atm}} = 260 \text{ L}$$

9. The information given is as follows:

   P = 5.00 atm T = 37^{o}C = 310 K The gas is SCl_2

The questions asks us to find the density.

The formula for gas density is $D = \frac{PM}{RT}$ where M is the molar mass.

Molar masas of SCl_2 is:

1 S x 32.06 = 32.06
2 Cl x 35.45 = <u>70.9</u>
102.96 g/mole

Plugging in we get:

$$D = \frac{PM}{RT} = \frac{(5.00\ \text{atm})(102.96\ \text{g/mole})}{(0.0821\ \frac{\text{L} \bullet \text{atm}}{\text{mole} \bullet \text{K}})(310\ \text{K})} = 20.2\ \text{g/L}$$

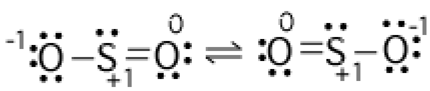
10. A. SO_2

Step 1 1 S x 6 = 6
2 O x 6 = <u>12</u>
18

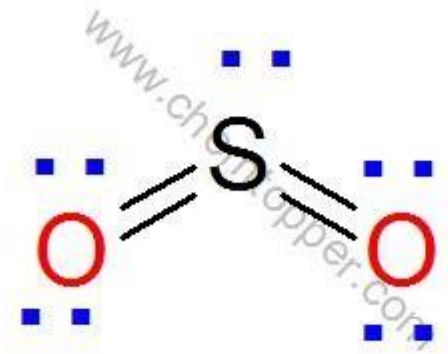
Step 2 1 s x 8 = 8
2 O x 8 = 16
24

Step 3 24 -18 = 6 BE

Step 4 6/2 = 3 bonds



From formal charges, the best structure is:



B. IO_3^-

Step 1 1 I x 7 = 7
3 O x 6 = 18
- sign = $\underline{1}$
26

Step 2 1 I x 8 = 8
3 O x 8 = $\underline{24}$
32

Step 3 32 - 26 = 6 BE Step 4 6/2 = 4 bonds

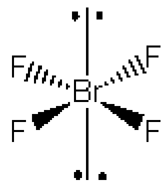
I could not find a Lewis structure for iodate so I will give a verbal description.
The Lewis structure without considering formal charges has three Iodine – oxygen single bonds and a lone pair on the central iodine atom.

C. BrF_4^-

Step 1 1 Br x 7 = 7
4 FO x 7 = 28
- sign = $\underline{1}$
36

Step 2 1 Br x 8 = 8
4 F x 8 = $\underline{32}$
40

Step 3 40 – 36 = 4 BE Step 4 4/2 = 2 bonds
Octet rules fail use exapanded octets.



D. HCN

Step 1 1 H x 1 = 1
1 C x 4 = 4
1 N x 5 = 5
10

Step 2 1 H x 2 = 2
1 C x 8 = 8
1 N x 8 = 8
18

Step 3 18 – 10 = 8 BE Step 4 8/2 = 4 bonds

H — C ≡ N:

11. The only species which exhibits resonance is SO_2.

12. The polar species all have lone pair on the central atom unsymmetrical bonding. They are: SO_2, IO_3^- and HCN

13. (12 points) Complete the following table (using your Lewis structures from question 11)

SUBSTANCE	ELECTRONIC GEOMETRY	MOLECULAR GEOMETRY	HYBRIDIZATION
SO_2	Trigonal Planar	Bent	sp^2
IO_3^-	Tetrahedral	Pyramidal	sp_3
BrF_4^-	Octahedral	Square Planar	sp^3d^2
HCN	Linear	Linear	Sp

14. Use the formulas: $c = \lambda \nu$ and $\Delta E = h\nu$ or $\Delta E = \frac{hc}{\lambda}$

Where h = 6.626 x 10^{-34} J sec and c = 3.0 x 10^{8} m/sec

Using $\Delta E = \frac{hc}{\lambda}$

We rearrange to solve for wavelength we get

$$\lambda = \frac{hc}{\Delta E} = \frac{(6.626 \text{ x } 10^{-34} \text{ J sec})(3 \text{ x } 10^{8} \frac{\text{m}}{\text{sec}})}{4.38 \text{ x } 10^{-18} \text{ J}} = 4.54 \text{ x } 10^{-8} \text{ m}$$

15. We need to use the equation q = (mass)(specific heat)(change in temperature)

We first need to find the mass of bromine using the volume and density.

Mass of bromine = 10 mL x $\frac{3.12 \text{ g}}{1 \text{ mL}} = 31.2 \text{ g of bromine}$

Now we can plug in:

Heat (q) = (31.2 g)(0.226 J/g^{o}C)((27.30 – 25.00^{o}C)
= (31.2 g)(0.226 J/g^{o}C)(2.30^{o}C) = 16.2 J

16. We are given a balanced equation and heat of formations for CO_2(g) and H_2O(*l*) and also ΔH_{rxn} and asked to find the heat of formation of $C_2H_5O_5$.

We use the formula: $\Delta H^{o}_{rx} = \sum \Delta H^{o}_{f}(products) - \sum \Delta H^{o}_{f}\text{ (reactants)}$

For this reaction:

ΔH_{rxn} = [(2 mole CO_2)(ΔH^{o}_{f} for CO_2) + (3 mole H_2O)(ΔH^{o}_{f} for H_2O(*l*)] –
[(1 mole)(ΔH_f for C_2H_5OH(*l*) + (3 mole)(ΔH_f for O_2(g)]

Since O_2 is in its standard state its ΔH_f is zero.

ΔH_{rxn} = -1368 (given in problem).

-1368 = [(2 mole)(-394 kJ?mole) + (3 mole)(-286 kJ/mole)] –

[(mole) (ΔH^o_f $C_2H_5OH(l)$) + (3 mole)(0)]

-1368 = [(-788 kJ) + (-858) kJ] – [(1 mole)(ΔH_{rxn} for $C_2H_5OH(l)$)]

-1368 = [-1646] -

[(1 mole)(ΔH_{rxn} for $C_2H_5OH(l)$)] = -1646 + 1368

[(1 mole)(ΔH_{rxn} for $C_2H_5OH(l)$)] = -278 kJ

Divide both sides by 1 mole to get

ΔH_{rxn} for $C_2H_5OH(l)$ = -278 kJ/mole

17. This is a Hess's law problem:

The reaction we want to find ΔH is $3\ Fe_2O_3(s) + CO(g) \rightarrow CO_2(g) + 2\ Fe_3O_4(s)$

The reactions we are given are:

$Fe_2O_3(s) + 3\ CO(g) \rightarrow 2\ Fe(s) + 3\ CO_2(g)$ $\Delta H = -28.0$ kJ

$3\ Fe(s) + 4\ CO_2(g) \rightarrow 4\ CO(g) + Fe_3O_4(s)$ $\Delta H = 12.5$ kJ

We need 3 mole of Fe_2O_3 on the reactant side in the reaction of interest

Reaction 1 above has Fe_2O_3 on the reactant side but only 1 mole of it, so we need to 2 multiply this reaction by 3 to get:

3 x ($Fe_2O_3(s) + 3\ CO(g) \rightarrow 2\ Fe(s) + 3\ CO_2(g)$ $\Delta H = -28.0$ kJ)

Remember we also need to multiply ΔH by 3.

Doing the multiplication we get:

$3\ Fe_2O_3(s) + 9\ CO(g) \rightarrow 6\ Fe(s) + 9\ CO_2(g)$ $\Delta H = 3(-28.0\text{ kJ}) = -84$ kJ

We need 2 mole of Fe_3O_4 on the product side in the reaction of interest

Reaction 2 above has Fe_3O_4 on the product side but only 1 mole of it, so we need to multiply this reaction by 2 to get:

2 x (3 Fe(s) + 4 CO_2(g) → 4 CO(g) + Fe_3O_4(s) $\Delta H = 12.5$ kJ)

Doing the multiplication we get:

6 Fe(s) + 8 CO_2(g) → 8 CO(g) + 2 Fe_3O_4(s) $\Delta H = 2(12.5 \text{ kJ}) = 25.0$ kJ

Now we can add the reactions together and see if we get the desired reaction.

3 Fe_2O_3(s) + 9 CO(g) → 6 Fe(s) + 9 CO_2(g) $\Delta H = -84$ kJ

6 Fe(s) + 8 CO_2(g) → 8 CO(g) + 2 Fe_3O_4(s) $\Delta H = 25.0$ kJ

Cancelling out like terms 6 mole Fe, 8 mole CO and 8 mole CO_2 we are left with:

3 Fe_2O_3(s) + CO(g) → CO_2(g) + 2 Fe_3O_4(s) $\Delta H = (-84 + 25)$

$\Delta H = -59.0$ kJ

18. In this problem the amount of heat released is directly proportional to the amount of material reactant. We can use the balanced equation and heat of reaction to create a conversion factor between ΔH_{rx} and the mole of any reactant or product in the balanced equation.

In this case $\Delta H = -192.8$ kJ = 2 mole SO_2 = 1 mole O_2 = 2 mole SO_3

For this problem, we need to create the conversions between ΔH and mole SO_3.

We get -192.8 kJ = 2 mole SO_3.

We need to convert 200 g SO_3 to mole and then use the above conversion factor to convert mole SO_3 to heat released.

$$200 \text{ g } SO_3 \times \frac{1 \text{ mole } SO_3}{80.0 \text{ g } SO_3} \times \frac{-192.8 \text{ kJ}}{2 \text{ mole } SO_3} = -241 \text{ kJ}$$

Since the problem told us heat was released, we do not need to include the negative sign so 241 kJ is the proper answer.