Name H-1 ______________________ Date ______________________

Dimensional Analysis I

Convert The Following Units (SHOW ALL WORK)

1) 5.5 kg = ________ cg

2) 25×10^5 nm = ________ m

3) 3.5 km/sec = ________ km/hr

4) 14 μg/l= ________mg/ml

5) 5.7×10^2 m^3 = ________ cm^3

6) 1.5×10^8 mg/cm^3 = ________ kg/mm^3

7) 7.0×10^{20} mm^3/sec= ________ m^3/day

8) *** 75 m^3 = (Hint: 1cm^3 = 1 ml) ________ l

H-2

Name ______________________ Date ______________________

Dimensional Analysis II

Convert The Following Units (SHOW ALL WORK)

1) 3 hrs = _______ seconds

2) 5 km = _______ pm

3) 3.5 cm/sec = _______ km/hr

4) 14 g/ml= _______ μg/cl

5) 14 cg/l= _______ mg/cl

6) 7.7 m^3 = _______ mm^3

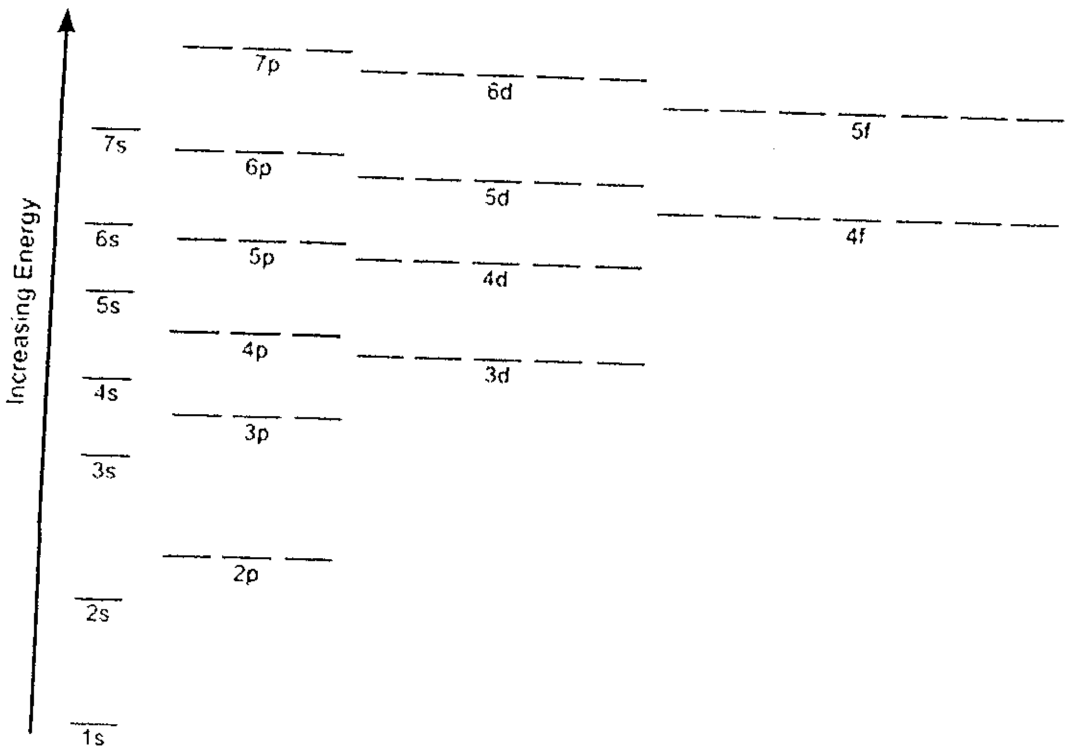
7) 15 mg/cm^3 = _______ kg/mm^3

8) 715 mm^3/sec= _______ m^3/day

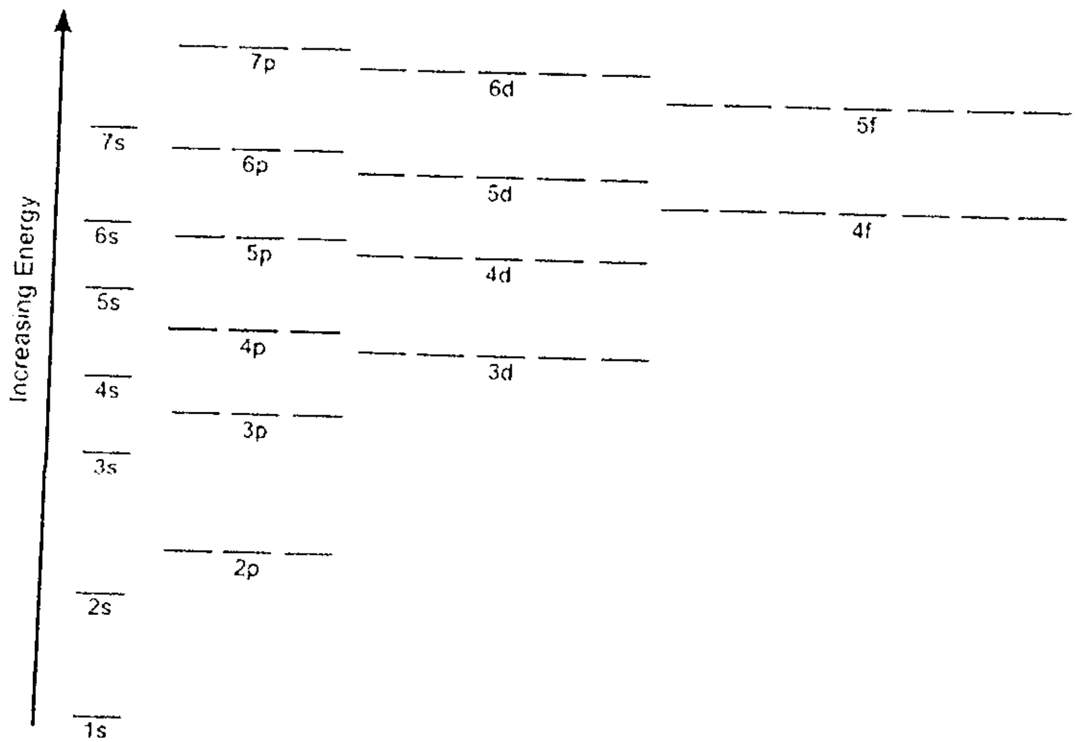
H-3

Orbital Diagram Worksheet

Element: ____ Electron Configuration: ______________________________



Element: ____ Electron Configuration: ______________________________



H-4

n ℓ m_ℓ m_s

1. State the four quantum numbers and the possible values they may have.

2. Name the orbitals described by the following quantum numbers

a. $n = 3, \ell = 0$
b. $n = 3, \ell = 1$
c. $n = 3, \ell = 2$
d. $n = 5, \ell = 0$

3. Give the n and ℓ values for the following orbitals

a. 1s
b. 3s
c. 2p
d. 4d
e. 5f

4. Place the following orbitals in order of increasing energy:

1s, 3s, 4s, 6s, 3d, 4f, 3p, 7s, 5d, 5p

5. What are the possible m_ℓ values for the following types of orbitals?

a. s
b. p
c. d
d. f

6. How many possible orbitals are there for n =

a. 4
b. 10

7. How many electrons can inhabit all of the n=4 orbitals?

Identify the element whose last electron would have the following four quantum numbers:

8. 3, 1, -1, +1/2
9. 4, 2, +1, +1/2
10. 6, 1, 0, -1/2
11. 4, 3, +3, -1/2
12. 2, 1, +1, -1/2

H-5

QUANTUM NUMBERS WORKSHEET **Name** ______________________________

1. State the four quantum numbers and the possible values they may have.

2. Name the orbitals described by the following quantum number
   a. n = 3, l = 0
   b. n = 3, l = 1
   c. n = 3, l = 2
   d. n = 5, λ = 0
3. Give the n and l values for the following orbitals
   a. 1s
   b. 3s
   c. 2p
   d. 4d
   e. 5f

4. Circle all of the following orbital destinations that are theoretically possible.
   a. 7s b. 1p c. 5d d. 2d e. 4f f. 5g g. 6i

5. Without referring to a text, periodic table or handout, deduce the maximum number of electrons that can occupy an:
   a. s orbital _____ b. the subshell of p orbitals _______ c. the subshell of d orbitals ____

   d. the subshell of f orbitals_______ e. the subshell of g orbitals_______

6. Circle all of the following electron configurations that are ruled out by the Pauli exclusion principle.
   a. $1s^2 2s^2 2p^7$ b. $1s^2 2s^2 2p^6 3s^3$ c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{12}$ d. $1s^2 2s^2 2p^6 3s^2 3p^6$

7. Explain why the following ground-state electron configurations are not possible:
   a. $1s^2 2s^3 2p^3$ b. $1s^2 2s^2 2p^3 3s^6$ c. $1s^2 2s^2 2p^7 3s^2 3p^8$ d. $1s^2 2s^2 2p^6 3s^2 3p^1 4s^2 3d^{14}$

8. Give two examples (i.e. list 2 elements that are examples) of:
   a. an atom with a half-filled subshell
   b. an atom with a completely filled outer shell
   c. an atom with its outer electrons occupying a half-filled subshell and a filled subshell.

9. Place the following orbitals in order of increasing energy:
   1s, 3s, 4s, 6s, 3d, 4f, 3p, 7s, 5d, 5p
10. What are the possible m_l values for the each of the following types of orbitals?
   a. s
   b. p
   c. d
   d. f

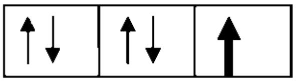
11. How many possible orbitals are there for n =
   a. 4
   b. 10

12. How many electrons can inhabit all of the n=4 orbitals?

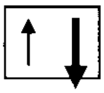
13. Fill in the blanks with the correct response:
  a. The number of orbitals with the quantum numbers n=3, l=2 and m_l = 0 is __________.
  b. The subshell with the quantum numbers n=4, l=2 is __________.
  c. The m_l values for a d orbital are ____________________________.
  d. The allowed values of l for the shell with n=2 are __________.
  e. The allowed values of l for the shell with n=4 are __________.
  f. The number of orbitals in a shell with n=3 is __________.
  g. The number of orbitals with n=3 and l=1 is __________.
  h. The maximum number of electrons with quantum numbers with n=3 and l=2 is __________.
  i. When n=2, l can be __________.
  j. When n=2, the possible values for m_l are __________.
  k. The number of electrons with n=4, l=1 is __________.
  l. The subshell with n=3 and l=1 is designated as the ___________ subshell.
  m. The lowest value of n for which a d subshell can occur is n=__________.

14. Which sets of quantum numbers are unacceptable? (Select a, b, c, or any combination)
  a. n=3, l= -2, m_l= 0, m_s= +½
  b. n=2, l= 2, m_l= -1, m_s= -½
  c. n=6, l= 2, m_l= -2, m_s= +½

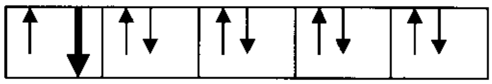
15. Write the values for the quantum numbers for the **bold** electron in the following diagrams:
  a. 3p orbitals



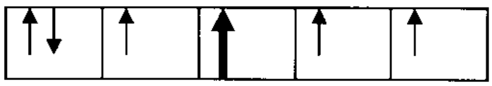
  b. 5s



  c. 4d orbitals



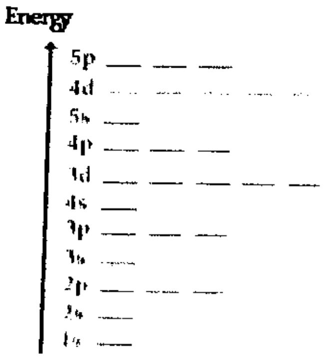
  d. 3d orbitals

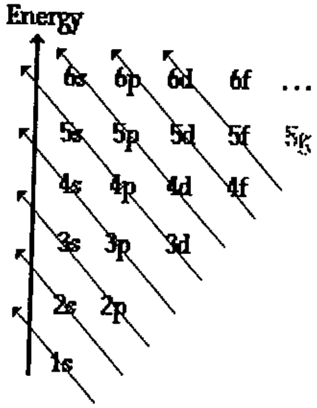


Electron Configurations and Orbital Diagrams

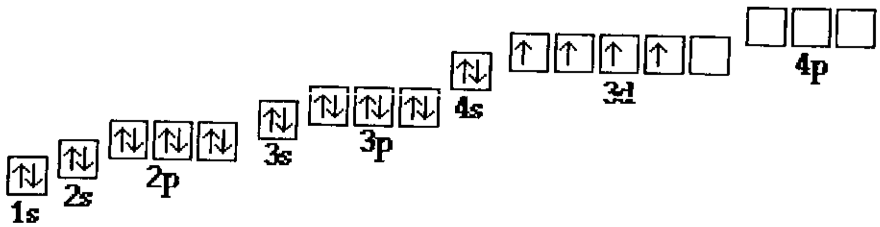
According to the Quantum Mechanical model of the atom, every electron of an atom is described by four quantum numbers. The quantum numbers describe the **orbitals** that the electrons are located in. Each orbital has a unique size (**n** value), shape (**l** value), and spatial orientation (**m_l** value). Each orbital can hold at most two electrons, with a *full* orbital having two electrons with a different spin direction (**m_s** value). The location of the electrons within the various orbitals is often expressed by **orbital diagrams** and **electron configuration** symbols.

Electrons fill the orbitals of an atom starting with the lowest energy level. Once each orbital at the same energy sublevel is filled, electrons begin filling the orbitals of the next energy sublevel. The ordering of the energy of the various orbital types is shown in the diagram at the left. It might be surprising to observe that the 4s orbitals of an atom have slightly lower energy than the 3d orbitals. While they are very close in energy, the 4s orbital is slightly lower in energy. The diagram at the right represents a convenient way of remembering the order in which the orbitals fill. Simply follow the arrows beginning with the lowest one. A final rule for filling orbitals with electrons is that each orbital at the same energy sublevel must have an electron before electrons begin pairing up inside the same orbital; this is known as Hund's rule.





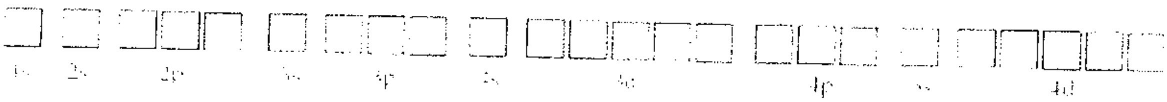
The diagram below represents the so-called orbital diagram for chromium. The 24 electrons of a chromium atom will fill each of the atomic orbitals in the manner shown.



An orbital diagram naturally leads to the writing of an electron configuration. The electron configuration for chromium is:

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$$

The orbital diagram above is formatted in such a manner as to place the various orbital types at different energy levels. A similar format that is used in the textbook (and serves to save space) is the format below in which the orbitals are listed in order of their energies but along the same line.



1. Show the orbital diagram for the following elements.

Be: 1s 2s 2p 3s 3p

C: 1s 2s 2p 3s 3p

F: 1s 2s 2p 3s 3p

Na: 1s 2s 2p 3s 3p

Al: 1s 2s 2p 3s 3p

S: 1s 2s 2p 3s 3p

2. Write electron configurations (showing where all the electrons are located) for the following. If necessary, use the orbital diagram to assist in the process.

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d

Element	# of Electrons	Electron Configuration
H		
He		
Li		
Be		
B		
C		
F		
Ne		
Na		
Mg		
Al		
S		
Ar		
K		
Ca		

3. Fill in the orbital diagram and write the electron configuration of the following:

a. Kr

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d

e- Configuration:

b. Br^-

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d

e- Configuration:

c. Rb^+

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d

e- Configuration:

d. Sc^{3+}

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d

e- Configuration:

e. Mn

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d

e- Configuration:

f. Cu^{2+}

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d

e- Configuration:

g. Fe^{2+}

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d

e- Configuration:

h. Rh

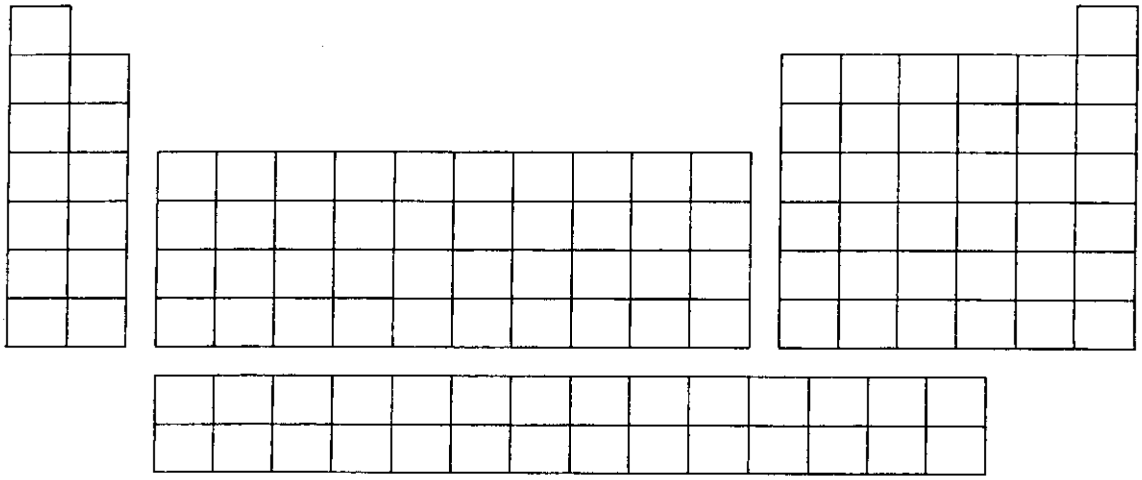
1s 2s 2p 3s 3p 4s 3d 4p 5s 4d

e- Configuration:

4. For the following groups of atoms and ions, circle those which have the same electron configuration:

a.	Al	B	F^-	Mg^{2+}	O^{2-}
b.	Ca^{2+}	Na^+	S^{2-}	Ba^{2+}	V^{3+}
c.	Se^{2-}	I^-	Sr^{2+}	Se	Ba
d.	Y^{3+}	Rb	Xe	Br^-	Kr

5. On the periodic table below, identify the s-block, p-block, d-block and f-block elements. Then make the connection between the various orbitals and the elements in the periodic table.



6. Use the noble gases to write **abbreviated electron configurations** for the following:

a. H ______________________________

b. Li ______________________________

c. O ______________________________

d. Mg ______________________________

e. Cl ______________________________

f. V ______________________________

g. EE ______________________________

h. Se^{2-} ______________________________

i. Cr^{2+} ______________________________

j. I^{-} ______________________________

k. Ag ______________________________

l. Sn^{2+} ______________________________

7. An ion of an isotope has a 2+ charge, an atomic mass of 56.9397 amu, 2 electrons at the n=4 energy level and 13 electrons at the n=3 energy level. Determine the ...

a. ... atomic number: ______________

b. ... mass number: ______________

c. ... total number of electrons: ______________

d. ... total number of s electrons: ______________

e. ... total number of p electrons: ______________

f. ... total number of d electrons: ______________

Honors Chemistry
Electron Arrangement/Quantum Numbers Worksheet

1. What is the frequency of light (EM radiation) if it is known to have a wavelength of 832 **cm**? (Frequency is in units of /sec). [Hint: make sure to use the proper units!]

   - What amount of energy is associated with a photon at this wavelength?

   - What amount of energy is associated with a mole of photons at this wavelength?

2. What element is it whose neutral, isolated atom has two valence electrons in the 5s subshell and four electrons in the 5p subshell?

3. How many energy levels (**shells**) are **either filled or partially filled** in the ground state configuration for the neodymiumium atom [Nd: Atomic #60]?

4. Based on our knowledge of the energy levels, what would be the predicted atomic number for the next noble gas (VIII-A) after radon (Neglect the hypothetical "g" subshell)?

5. How many **valence** electrons are there in an arsenic atom?

6. Which of the following is an electron configuration of an atom in the excited state? (May be more than 1)

   A) $1s^2\ 2s^1$; B) $1s^2\ 2s^2\ 2p^1$; **C)** $1s^2\ 2s^1\ 3d^6$; **D)** $1s^2\ 2s^1\ 2p^2$; **E)** [Ar] $3d^8$.

7. How many sublevels are either filled or partially filled in an atom of iodine in the ground state?

8. What is the total number of electrons in an atom if the **fourth shell** has been completely filled? Ignore any abnormalities in electron configuration, but remember the order as indicated by the Diagonal Rule and Aufbau Principle.

9. Which quantum number value indicates the most probable distance of the electron from the nucleus of an atom?

10. The **f** sublevel contains how many electrons? How many orbitals?

11. What is the next subshell to be filled immediately after **6s**?

12. How many **filled 3p orbitals** are there in a neutral atom of cobalt?

13. What type of subshell is being filled to make the transition metals?

14. What type of subshell is being filled to make the rare-earth metals, or inner transition metals?

15. According to Bohr, the total energy of an electron in an atom is quantized. This means that:
A) an atom's electrons can possess only certain specific amounts of energy;
B) the electrons travel about the nucleus in orbits;
C) the electrons sometimes behave as waves;
D) the position and momentum of an electron cannot be known simultaneously.

16. Which of the following would be expected to have the **shortest** wavelength?
A) gamma; B) TV; C) visible; D) X rays; E) IR; F) microwave

17. The orbitals within each principal quantum number can be designated by what values?

18. The subshells within each principal quantum number can be designated by what values?

19. What is the maximum number of electrons that can fill any one of the seven "f" orbitals?

20. What rule states that no two electrons in the same atom can have the same set of four quantum numbers?

21. What sublevel does the **forty-third** electron of a neutral polonium atom occupy in the ground state?

GIVEN THE FOLLOWING ELECTRON CONFIGURATION:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}6p^4$

Answer the following questions referring to the above configuration:

22. In which **sublevel** are the electrons of highest energy located?

23. What is the correct chemical symbol for element above?

24. What is the number of electron pairs in the 5d orbital?

25. This element would most likely be classified as what type of element?

26. The correct number of outer shell electrons for this element is how many?

27. This element is radioactive - true or false?

28. How many sublevels are either filled or partially filled?

29. An orbital is labeled by the magnetic quantum number, m = +2. This could not be found in which subshell(s)?

30. How many electrons can are there in a neutral atom in which the last electron to enter has just filled the 5p orbital?

31. In a chemical reaction, the removal, or movement, of what type of electrons is associated with the color of the salts of the transition metals?

32. The idea of electron energy levels in the hydrogen atom was developed by what scientist?

33. What scientist made a major contribution to science with his work on the photoelectric effect?

34. What does the orbital quantum number (m) indicate?

35. The **L** shell corresponds to which principal quantum number?

36. Which of the following is(are) a valid set of quantum numbers for one of the electrons in the ground state of a cerium atom? (Not necessarily the last electron.)

	n	l	m	s
A)	1	0	+1	+1/2
B)	5	3	-1	+1/2
C)	2	2	-1	-1/2
D)	1	1	0	+1/2
E)	6	0	0	-1/2
F)	6	2	-2	+1/2
G)	4	3	0	+1/2
H)	5	0	0	-1/2

37. What is the number of possible orbitals in the **d** subshell?

38. If **n** represents the principal quantum number of an energy level, what is the number of orbitals in that energy level is equal to? (in a formula form)

39. The maximum number of electrons that may occupy a **p** sublevel is......

40. Which of the following elements in the fourth series (elements #19 - 36) whose electron configuration appears to be irregular because of the stability of a <u>completely filled</u> sublevel?

41. What element has two filled and one half-filled 5p orbitals in its neutral atom?

42. For the Lewis-dot diagram of a neutral silicon atom, how many dots must surround the symbol?

43. What rule states that electrons in any subshell tend to stay unpaired as long as possible?

44. What element has five half-filled 4d orbitals in its +2 cationic form?

45. What is the maximum number of **p** electrons (**total**) available in a neutral atom of iodine?

46. If a neutral atom loses electrons, what does it become?

47. With the exception of helium, a noble gas occurs with the completion of what subshell?

48. How many electrons are in the outer shell of a neutral strontium atom?

49. Which of the following ions is isoelectronic to a neutral atom of krypton?

A) S^{-2}; B) Se^{-2}; C) Sr^{2+}; D) Sb^{-3}; **E)** both B and C.

50. What is the amount of energy released if a photon with a wavelength of 4.88×10^{-14} meters is released as an electron drops to ground state?

51. If a **mole** of electrons releases 7.73×10^{-3} joules of energy in the form of photons as they drop to their ground state, what is the wavelength of **one** of these photons?

52. A cation with a +3 charge has the following electron configuration: [Ar] What element is it?

53. What type of electromagnetic radiation has a **shorter** wavelength than **gamma** radiation?

54. Give the **complete** electron configuration for the nickel +2 ion.

55. What is the term for a packet of energy?

56. How many filled "p" orbitals are there in a neutral atom of tellurium?

57. Describe the shapes of each of the orbital types.

58. Give a characteristic that distinguishes absorption spectrum from an emission spectrum.

59. What is the significance of the Heisenberg Uncertainty Principle?

60. What is the significance of de Broglie's wave equation?

61. Explain how incandescent lights work. How do fluorescent lights work?
Which is more efficient? Why?

62. What importance is a flame test? Who is credited with the development of this test?

63. What color is imparted when **sodium** is added to a flame?
What color is **copper**? **Barium**? **Lithium**?

Periodic Table of the Elements

KEY

Atomic Mass →	12.011	−4 +2 +4 ← Selected Oxidation States
Symbol →	**C**	
Atomic Number →	6	
Electron Configuration →	2-4	

Relative atomic masses are based on $^{12}C = 12$ (exact)

Note: Numbers in parentheses are mass numbers of the most stable or common isotope.

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1.00794 **H** 1 1 +1 −1																	4.00260 **He** 2 2 0
2	6.941 **Li** 3 2-1 +1	9.01218 **Be** 4 2-2 +2											10.81 **B** 5 2-3 +3	12.011 **C** 6 2-4 −4 +2 +4	14.0067 **N** 7 2-5 −3 −2 −1 +1 +2 +3 +4 +5	15.9994 **O** 8 2-6 −2	18.9984 **F** 9 2-7 −1	20.180 **Ne** 10 2-8 0
3	22.98977 **Na** 11 2-8-1 +1	24.305 **Mg** 12 2-8-2 +2											26.98154 **Al** 13 2-8-3 +3	28.0855 **Si** 14 2-8-4 −4 +2 +4	30.97376 **P** 15 2-8-5 −3 +3 +5	32.065 **S** 16 2-8-6 −2 +4 +6	35.453 **Cl** 17 2-8-7 −1 +1 +5 +7	39.948 **Ar** 18 2-8-8 0
4	39.0983 **K** 19 2-8-8-1 +1	40.08 **Ca** 20 2-8-8-2 +2	44.9559 **Sc** 21 2-8-9-2 +3	47.867 **Ti** 22 2-8-10-2 +2 +3 +4	50.9415 **V** 23 2-8-11-2 +2 +3 +4 +5	51.996 **Cr** 24 2-8-13-1 +2 +3 +6	54.9380 **Mn** 25 2-8-13-2 +2 +3 +4 +7	55.845 **Fe** 26 2-8-14-2 +2 +3	58.9332 **Co** 27 2-8-15-2 +2 +3	58.693 **Ni** 28 2-8-16-2 +2 +3	63.546 **Cu** 29 2-8-18-1 +1 +2	65.409 **Zn** 30 2-8-18-2 +2	69.723 **Ga** 31 2-8-18-3 +3	72.64 **Ge** 32 2-8-18-4 −4 +2 +4	74.9216 **As** 33 2-8-18-5 −3 +3 +5	78.96 **Se** 34 2-8-18-6 −2 +4 +6	79.904 **Br** 35 2-8-18-7 −1 +1 +5	83.798 **Kr** 36 2-8-18-8 0 +2
5	85.4678 **Rb** 37 2-8-18-8-1 +1	87.62 **Sr** 38 2-8-18-8-2 +2	88.9059 **Y** 39 2-8-18-9-2 +3	91.224 **Zr** 40 2-8-18-10-2 +4	92.9064 **Nb** 41 2-8-18-12-1 +3 +5	95.94 **Mo** 42 2-8-18-13-1 +6	(98) **Tc** 43 2-8-18-13-2 +4 +6 +7	101.07 **Ru** 44 2-8-18-15-1 +3	102.906 **Rh** 45 2-8-18-16-1 +3	106.42 **Pd** 46 2-8-18-18 +2 +4	107.868 **Ag** 47 2-8-18-18-1 +1	112.41 **Cd** 48 2-8-18-18-2 +2	114.818 **In** 49 2-8-18-18-3 +3	118.71 **Sn** 50 2-8-18-18-4 +2 +4	121.760 **Sb** 51 2-8-18-18-5 −3 +3 +5	127.60 **Te** 52 2-8-18-18-6 −2 +4 +6	126.904 **I** 53 2-8-18-18-7 −1 +1 +5 +7	131.29 **Xe** 54 2-8-18-18-8 0 +2 +4 +6
6	132.905 **Cs** 55 2-8-18-18-8-1 +1	137.33 **Ba** 56 2-8-18-18-8-2 +2	138.9055 **La** 57 2-8-18-18-9-2 +3	178.49 **Hf** 72 *18-32-10-2 +4	180.948 **Ta** 73 -18-32-11-2 +5	183.84 **W** 74 -18-32-12-2 +6	186.207 **Re** 75 -18-32-13-2 +4 +6 +7	190.23 **Os** 76 -18-32-14-2 +3 +4	192.217 **Ir** 77 -18-32-15-2 +3 +4	195.08 **Pt** 78 -18-32-17-1 +2 +4	196.967 **Au** 79 -18-32-18-1 +1 +3	200.59 **Hg** 80 -18-32-18-2 +1 +2	204.383 **Tl** 81 -18-32-18-3 +1 +3	207.2 **Pb** 82 -18-32-18-4 +2 +4	208.980 **Bi** 83 -18-32-18-5 +3 +5	(209) **Po** 84 -18-32-18-6 +2 +4	(210) **At** 85 -18-32-18-7	(222) **Rn** 86 -18-32-18-8 0
7	(223) **Fr** 87 -18-32-18-8-1 +1	(226) **Ra** 88 -18-32-18-8-2 +2	(227) **Ac** 89 -18-32-18-9-2 +3	(261) **Rf** 104 +4	(262) **Db** 105	(266) **Sg** 106	(272) **Bh** 107	(277) **Hs** 108	(276) **Mt** 109	(281) **Ds** 110	(280) **Rg** 111	(285) **Cn** 112	(284) **Uut** 113**	(289) **Uuq** 114	(288) **Uup** 115	(292) **Uuh** 116	(?) **Uus** 117	(294) **Uuo** 118

57–71	140.116 **Ce** 58 +3 +4	140.908 **Pr** 59 +3	144.24 **Nd** 60 +3	(145) **Pm** 61 +3	150.36 **Sm** 62 +2 +3	151.964 **Eu** 63 +2 +3	157.25 **Gd** 64 +3	158.925 **Tb** 65 +3	162.500 **Dy** 66 +3	164.930 **Ho** 67 +3	167.259 **Er** 68 +3	168.934 **Tm** 69 +3	173.04 **Yb** 70 +2 +3	174.9668 **Lu** 71 +3
89–103	232.038 **Th** 90 +4	231.036 **Pa** 91 +4 +5	238.029 **U** 92 +3 +4 +5 +6	(237) **Np** 93 +3 +4 +5 +6	(244) **Pu** 94 +3 +4 +5 +6	(243) **Am** 95 +3 +4 +5 +6	(247) **Cm** 96 +3	(247) **Bk** 97 +3 +4	(251) **Cf** 98 +3	(252) **Es** 99 +3	(257) **Fm** 100 +3	(258) **Md** 101 +2 +3	(259) **No** 102 +2 +3	(262) **Lr** 103 +3

*denotes the presence of (2-8-) for elements 72 and above

**The systematic names and symbols for elements of atomic numbers 113 and above will be used until the approval of trivial names by IUPAC.

Source: *CRC Handbook of Chemistry and Physics*, 91st ed., 2010–2011, CRC Press

Hon 7 (H)

Solutions

1) Write the balanced equation for the reaction that occurs when iron (II) chloride is mixed with sodium phosphate forming iron (II) phosphate and sodium chloride.

$FeCl_2$ + Na_3PO_4 → $Fe_3(PO_4)_2$ + NaCl

2) If 23 grams of iron (II) chloride reacts with 41 grams of sodium phosphate, what is the limiting reagent? How much sodium chloride can be formed?

$$23\text{ g }FeCl_2 \times \frac{1\text{ mole }FeCl_2}{126.75\text{ g }FeCl_2} \times \frac{2\text{ mole }Na_3PO_4}{3\text{ mole }FeCl_2} \times \frac{163.94\text{ g }Na_3PO_4}{1\text{ mole }Na_3PO_4} =$$

= 20. g Na_3PO_4 Since we have 41 g Na_3PO_4, $FeCl_2$ is the limiting reagent.

$$23\text{ g }FeCl_2 \times \frac{1\text{ mole }FeCl_2}{126.75\text{ g }FeCl_2} \times \frac{6\text{ mole NaCl}}{3\text{ mole }FeCl_2} \times \frac{58.44\text{ g NaCl}}{1\text{ mole NaCl}} =$$

=

3) How much of the excess reagent remains when this reaction has gone to completion?

41 g Na_3PO_4 - 20. g Na_3PO_4 =

4) If 16.1 grams of sodium chloride are formed in the reaction, what is the percent yield of this reaction?

$$\frac{16.1\text{ g NaCl}}{21\text{ g NaCl}} \times 100 =$$

Solutions

1) Write a balanced equation for the reaction of tin (IV) phosphate with sodium carbonate to make tin (IV) carbonate and sodium phosphate.

2) If 36 grams of tin (IV) phosphate is mixed with an excess of sodium carbonate, how many grams of tin (IV) carbonate will form?

$$36\ g\ Sn_3(PO_4)_4 \times \frac{1\ mole\ Sn_3(PO_4)_4}{736\ g\ Sn_3(PO_4)_4} \times \frac{3\ mole\ Sn(CO_3)_2}{1\ mole\ Sn_3(PO_4)_4} \times \frac{238.73\ g\ Sn(CO_3)_2}{1\ mole\ Sn(CO_3)_2} =$$

$$=$$

3) If 29.8 grams of tin (IV) carbonate are actually formed when this reaction goes to completion, what is the percent yield?

$$\frac{29.8\ g\ Sn(CO_3)_2}{35\ g\ Sn(CO_3)_2} \times 100 =$$

4) If 7.3 grams of sodium carbonate are used in the reaction and the result a 74.0% yield, how many grams of sodium phosphate will be formed?

$$7.3\ g\ Na_2CO_3 \times \frac{1\ mole\ Na_2CO_3}{105.99\ g\ Na_2CO_3} \times \frac{4\ mole\ Na_3PO_4}{1\ mole\ Na_2CO_3} \times \frac{163.94\ g\ Na_3PO_4}{1\ mole\ Na_3PO_4} =$$

$$=$$

$(7.5\ g\ Na_3PO_4)\ (0.74) =$

Limiting Reagent Worksheet

W 324

1) Write the balanced equation for the reaction that occurs when iron (II) chloride is mixed with sodium phosphate forming iron (II) phosphate and sodium chloride.

2) If 23 grams of iron (II) chloride reacts with 41 grams of sodium phosphate, what is the limiting reagent? How much sodium chloride can be formed?

3) How much of the excess reagent remains when this reaction has gone to completion?

4) If 16.1 grams of sodium chloride are formed in the reaction, what is the percent yield of this reaction?

H-20

Name ____________________________ Date ___________

Mass – Volume Problems

Solve each of the following problems for volume

1) 1 mole of $H_{2(g)}$	8) 5 grams of $H_{2(g)}$
2) 3.20 moles of $O_{2(g)}$	9) 100 grams of $O_{2(g)}$
3) .750 moles of $N_{2(g)}$	10) 28 grams of $N_{2(g)}$
4) 0.50 moles of $NH_{3(g)}$	11) 50 grams of $NH_{3(g)}$
5) 5 moles of $H_2SO_{4(g)}$	12) 150 grams of $H_2SO_{4(g)}$
6) .3 Moles of $C_6H_{12}O_{6(g)}$	13) 500 grams of $C_6H_{12}O_{6(g)}$
7) 1.75 moles of $Na_2CO_{3(g)}$	14) 15 grams of $Na_2CO_{3(g)}$

Name ______________________ Date ______________

Balancing & Mass - Mass

1. Balance
2. Convert: Mass → Moles
3. Convert: Moles → Moles
4. Convert: Moles → Mass

Reaction:

_____ $NaBr$ + _____ $Ca(OH)_2$ → ____ $CaBr_2$ + _____ $NaOH$

1) How many grams of sodium hydroxide will be produced using 75 grams of sodium bromide?

Answer

_____ NH_3+ _____ H_2SO_4 → _____ $(NH_4)_2SO_4$

2) How many grams of dihydrogen sulfate will be produced using 50 grams of nitrogen trihydride?

Answer

_____ $C_5H_{12}O$ + _____ O_2 → _____ CO_2 + _____ H_2O

3) How many grams of Oxygen are needed to produce 100 grams of carbon dioxide?

Answer

_____ Pb + _____ H_3PO_4 → _____ H_2 + _____ $Pb_3(PO_4)_2$

4) How many grams of lead will be need to react completely with 150 grams of trihydrogen phosphate?

Answer

7

_____ Li_3N + _____ NH_4NO_3 → ____ $LiNO_3$ + ____ $(NH_4)_3N$

5) How many grams of ammonium nitride are needed to react completely with 15 grams of lithium nitride?

Answer

_____ HBr + ____ $Al(OH)_3$ → ____ H_2O + ____ $AlBr_3$

6) How many grams of aluminum hydroxide will react with 50 grams of hydrogen monobromide?

Answer

_____ $C_6H_{12}O_6$ → ____ CH_4 + _____ CO_2

7) How many grams carbon dioxide will 500 grams of glucose produce?

Answer

_____ Mg + _____ $Ca(OH)_2$ → ____ Ca + _____ $Mg(OH)_2$

8) How many grams of magnesium are needed to react completely with 17 grams of calcium hydroxide?

Answer

_____ CH_4 + _____ O_2 → ____ CO_2 + _____ H_2O

9) How many grams of water will be produced using 300 grams of carbon tetrahydride?

Answer

H-23

Name: ______________________ Date: ______________

Period: ______________ Mr. Roderick

Stoichiometry Worksheet

Balance the following equations:

1) ___ N_2 + ___ F_2 → ___ NF_3

2) ___ C_6H_{10} + ___ O_2 → ___ CO_2 + ___ H_2O

3) ___ HBr + ___ $KHCO_3$ → ___ H_2O + ___ KBr + ___ CO_2

4) ___ $GaBr_3$ + ___ Na_2SO_3 → ___ $Ga_2(SO_3)_3$ + ___ NaBr

5) ___ SnO + ___ NF_3 → ___ SnF_2 + ___ N_2O_3

Using the equation from problem 2 above, answer the following questions:

6) If I do this reaction with 35 grams of C_6H_{10} and an excess of oxygen, how many grams of carbon dioxide will be formed?

7) Using the following equation:

$$2\ NaOH + H_2SO_4 \rightarrow 2\ H_2O + Na_2SO_4$$

How many grams of sodium sulfate will be formed if you start with 200 grams of sodium hydroxide and you have an excess of sulfuric acid?

H-24

Name: ______________________ Date: ______________

Period: ______________ Mr. Roderick

8) Using the following equation:

$$Pb(SO_4)_2 + 4\ LiNO_3 \rightarrow Pb(NO_3)_4 + 2\ Li_2SO_4$$

How many grams of lithium nitrate will be needed to make 250 grams of lithium sulfate, assuming that you have an adequate amount of lead (IV) sulfate to do the reaction?

Unit 3 (H)
H-25

Name ______________________________ Date ________________

Mass/Mass Problems

Honors

1. __$HCl_{(g)}$ + __$CaCO_{3(s)}$ → __$CaCl_{2(s)}$ + __$H_2O_{(g)}$ + __$CO_{2(g)}$

a. How many grams of calcium carbonate can be produced by the reaction of 150 grams of hydrogen chloride?

b. How many grams of calcium carbonate are needed to produce 75 grams of carbon dioxide?

c. How many liters of dihydrogen monoxide gas can be produced by 54 grams of hydrogen chloride?

d. How many liters of carbon dioxide will be produced by the reaction of 5 moles of calcium carbonate?

2. __$AgNO_3$ + __$MgCl_2$ → __$AgCl$ + __$Mg(NO_3)_2$

a. Silver nitrate reacts wil magnesium chloride to produce silver chloride. How much silver chloride can be produced if you have 11 moles of silver nitrate?

b. How much magnesium chloride is needed to produce 5 moles of magnesium nitrate?

c. How many grams of magnesium chloride are needed to react with 3 moles of silver nitrate?

d. If given 7 moles of magnesium chloride react with 6 moles of silver nitrate. What is the maximum amount of grams of silver chloride that can be produced? Which reagent runs out first? How much will remain unreacted?

3. __K_3PO_4 + __$Al(NO_3)_3$ → __KNO_3 + __$AlPO_4$

a. How many moles of potassium nitrate are produced from two moles of potassium phosphate?

b. How many grams of silver nitrate are needed to produce 15 moles of magnesium nitrate?

c. How many grams of aluminum phosphate are produced from 500 grams of potassium phosphate?

4. __$AlBr_3$ + __K_2SO_4 → __KBr + __$Al_2(SO_4)_3$

a. How much potassium bromide can be produced using seven moles of aluminum bromide?

b. How many moles of potassium sulfate are needed to produce 14 moles of aluminum sulfate??

c. How many grams of aluminum sulfate can be produced from 10 moles aluminum bromide and 10 moles of potassium sulfate? Which reactant will run out first? How many grams will remain unreacted?

H-27

Name ____________________________ Date _______

Limiting Reagents I

1. At high temperatures, sulfur combines with iron to form the brown-black iron (II) sulfide:

$$\_\_Fe_{(s)} + \_\_S_{(l)} \rightarrow \_\_FeS_{(s)}$$

   a. If 7.62 g of Fe are allowed to react with 8.67 g of S. What is the limiting reagent, and what is the reactant in excess?

   b. Calculate the mass of FeS formed.

2. Arcylonitrile, C_3H_3N, is the starting material for the production of a kind of synthetic fiber acrylics) and can be made from propylene, C_3H_6, by reaction with nitric oxide, NO, as follows:

$$\_\_C_3H_{6\,(g)} + \_\_NO_{(g)} \rightarrow \_\_C_3H_3N_{(s)} + \_\_H_2O_{(l)} + \_\_N_{2\,(g)}$$

   What mass of C_3H_3N can be made when 21.6 g of C_3H_6 react with 21.6 g of nitric oxide?

3. Formic acid, $HCHO_2$, burns in oxygen to form carbon dioxide and water as follows:

$$\_\_HCHO_{2\,(aq)} + \_\_O_{2\,(g)} \rightarrow \_\_CO_{2\,(g)} + \_\_H_2O_{(l)}$$

   If a 3.15-g sample of formic acid was burned in 2.0 L of oxygen, what volume of carbon dioxide would be produced? (Assume the reaction occurs at standard temperature and pressure, STP.)

4. Zinc metal reacts with hydrochloric acid to produce zinc chloride and hydrogen gas.

$$\_\_Zn_{(s)} + \_\_HCl_{(aq)} \rightarrow \_\_ZnCl_{2\,(aq)} + \_\_H_{2\,(g)}$$

A 3.5 g sample of zinc metal is allowed to react with 2.5 g of hydrochloric acid.

Complete the following table:

Reactants/products	Zn (grams)	HCl (grams)	$ZnCl_2$ (grams)	H_2 (L)
Before reaction				
After reaction	1.26 g			

5.

$$\_\_MnO_2 + \_\_HCl \rightarrow \_\_MnCl_2 + \_\_Cl_2 + \_\_H_2O$$

If 0.45 mols of MnO_2 can react with 48.2 g of HCl, how many grams of Cl_2 could be produced?

6. One of the components of the fuel mixture on the Apollo lunar module involved a reaction with hydrazine, N_2H_4, and dinitrogen tetraoxide, N_2O_4 If the equation for this reaction is

$$\_\_N_2H_{4\,(l)} + \_\_N_2O_{4\,(g)} \rightarrow \_\_N_{2\,(g)} + \_\_H_2O_{(g)}$$

What volume of N_2 gas (measured at STP) would result from the reaction of 1500 kg of hydrazine and 1000 kg of N_2O_4?

7. Chlorine gas reacts with silica, SiO_2, and carbon to give silicon tetrachloride and carbon monoxide.

$$\_\_Cl_{2\,(g)} + \_\_SiO_{2\,(s)} + \_\_C_{\,(s)} \rightarrow \_\_SiCl_{4\,(l)} + \_\_CO_{\,(g)}$$

How much CO gas can be produced from 15.0 g of silica?

8. When iron (II) hydroxide is mixed with phosphoric acid, iron (II) phosphate precipitate results.

$$\_\_Fe(OH)_{2\,(aq)} + \_\_H_3PO_{4\,(aq)} \rightarrow \_\_Fe_3(PO_4)_{2\,(s)} + \_\_H_2O_{\,(l)}$$

a. If 3.20 g of $Fe(OH)_2$ is treated with 2.50 g of phosphoric acid, what is the limiting reagent and what is the reactant in excess?

b. How many grams of $Fe_3(PO_4)_2$ precipitate can be formed?

c. If 1.99 g of $Fe_3(PO_4)_2$ is actually obtained, what is the percent yield?

9. In an experiment, 3.25 g of NH_3 are allowed to react with 3.50 g of O_2.

$$\_\_NH_3 + \_\_O_2 \rightarrow \_\_NO + \_\_H_2O$$

a. Which reactant is the limiting reagent?

b. How many grams of NO are formed?

c. How much of the excess reactant remains after the reaction?

10. If 4.95 g of ethylene (C_2H_4) are combusted with 3.25 g of oxygen.

$$\_\_C_2H_{4\,(g)} + \_\_O_{2\,(g)} \rightarrow \_\_H_2O_{(g)} + \_\_CO_{2\,(g)}$$

a. What is the limiting reagent?

b. How many grams of CO_2 are formed?

H-31

Name ______________________________ Date ______

Limiting Reagents II

1. For the reaction:

$$\_\_K + \_\_O_2 \rightarrow \_\_K_2O$$

a. If 0.50 g of K is reacted 0.10 g of O_2, what is the limiting reagent?

c. How many grams of K_2O will be produced from the reaction?

2. For the reaction:

$$\_\_Na_2O + \_\_H_2O \rightarrow \_\_NaOH$$

a. What weight of NaOH could be made from 12.4 g of Na_2O and 42.1 g of H_2O?

b. What would be the limiting reagent if 100 g each of Na_2O and H_2O were allowed to react?

3. For the reaction:

$$\_\_C + \_\_H_2 \rightarrow \_\_CH_4$$

a. How many moles of CH_4 can be made from 7.0 moles of H_2 and 5.0 moles of C?

b. What weight of CH_4 will be made when 10.0 g of H_2 reacts with 5.0 g of C?

4. Heating zinc sulfide in the presence of oxygen yields the following:

$$\_\_ZnS + \_\_O_2 \rightarrow \_\_ZnO + \_\_SO_2$$

a. If 1.72 mol of ZnS is heated in the presence of 3.04 mol of O_2, which reactant will be used up first?

b. What is the maximum amount of zinc oxide that can be produce?

5. Aluminum reacts with oxygen in a synthesis reaction.

$$\_\_Al_{(s)} + \_\_O_{2\,(g)} \rightarrow \_\_Al_2O_{3\,(s)}$$

a. Which reactant is limiting if 0.32 mol Al and 0.26 mol O2 are available?

b. How many moles of aluminum oxide are formed from the reaction of 6.38 mol of oxygen and 9.15 of Al?

c. If 3.17 g of Al and 2.55 g of O_2 are available, which reactant is limiting?

6. In the production of copper from ore containing copper (II) sulfide, the ore is first roasted to change it to the oxide according to the following equation:

$$\_\_CuS + \_\_O_2 \rightarrow \_\_CuO + \_\_SO_2$$

a. If 100 g of CuS and 56 g O_2 are available, which reactant is limiting?

b. What mass of CuO can be formed from the reaction of 18.7 g of CuS and 12.0 g of O_2?

7. Consider the following reaction:

$$\_\_NH_4NO_3 + \_\_Na_3PO_4 \rightarrow \_\_(NH_4)_3PO_4 + \_\_NaNO_3$$

a. If 50 grams of ammonium nitrate and 30 grams of sodium phosphate are available, which is the limiting reagent?

b. How much ammonium phosphate can be produced?

8. Consider the following reaction:

$$3\ CaCO_3 + 2\ FePO_4 \rightarrow Ca_3(PO_4)_2 + Fe_2(CO_3)_3$$

a. How many grams of iron (III) carbonate can be produce from 100 grams of calcium carbonate and 100 g of iron (III) phosphate?

b. What was the limiting reagent?

9. A double replacement reaction occurs between copper (II) sulfate and iron (III) nitrate:

$$\_\_CuSO_4 + \_\_Fe(NO_3)_3 \rightarrow \_\_Fe_2(SO_4)_3 + \_\_Cu(NO_3)_2$$

a. If you place 0.092 mol of iron (III) nitrate in a solution containing 0.158 mol of $CuSO_4$. What is the limiting reactant?

b. How many moles of $Cu(NO_3)_2$ will be formed?

10. In the reaction barium carbonate reacts with nitric acid to form barium nitrate, carbon dioxide and water.

$$\_\_Ba(CO_3) + \_\_HNO_3 \rightarrow \_\_Ba(NO_3)_2 + \_\_CO_2 + \_\_H_2O$$

a. What mass of barium nitrate can be formed by combining 55 g of barium carbonate with 26 g of nitric acid.

b. How many <u>liters</u> of $CO_{2\,(g)}$ will be produced if 15 moles of barium carbonate reacts with 100 grams of nitric acid?

Name ______________________________ Date ______

Limiting Reagents II

1. For the reaction:

$$\_\_K + \_\_O_2 \rightarrow \_\_K_2O$$

a. If 0.50 g of K is reacted 0.10 g of O_2, what is the limiting reagent?

c. How many grams of K_2O will be produced from the reaction?

2. For the reaction:

$$\_\_Na_2O + \_\_H_2O \rightarrow \_\_NaOH$$

What weight of NaOH could be made from 12.4 g of Na_2O and 42.1 g of H_2O?

b. What would be the limiting reagent if 100 g each of Na_2O and H_2O were allowed to react?

3. For the reaction:

$$\_\_C + \_\_H_2 \rightarrow \_\_CH_4$$

a. How many moles of CH_4 can be made from 7.0 moles of H_2 and 5.0 moles of C?

b. What weight of CH_4 will be made when 10.0 g of H_2 reacts with 5.0 g of C?

4. Heating zinc sulfide in the presence of oxygen yields the following:

$$\_\_ZnS + \_\_O_2 \rightarrow \_\_ZnO + \_\_SO_2$$

a. If 1.72 mol of ZnS is heated in the presence of 3.04 mol of O_2, which reactant will be used up first?

b. What is the maximum amount of zinc oxide that can be produce?

5. Aluminum reacts with oxygen in a synthesis reaction.

$$\_\_Al_{(s)} + \_\_O_{2\,(g)} \rightarrow \_\_Al_2O_{3\,(s)}$$

a. Which reactant is limiting if 0.32 mol Al and 0.26 mol O2 are available?

b. How many moles of aluminum oxide are formed from the reaction of 6.38 mol of oxygen and 9.15 of Al?

c. If 3.17 g of Al and 2.55 g of O_2 are available, which reactant is limiting?

6. In the production of copper from ore containing copper (II) sulfide, the ore is first roasted to change it to the oxide according to the following equation:

$$\_\_CuS + \_\_O_2 \rightarrow \_\_CuO + \_\_SO_2$$

a. If 100 g of CuS and 56 g O_2 are available, which reactant is limiting?

b. What mass of CuO can be formed from the reaction of 18.7 g of CuS and 12.0 g of O_2?

7. Consider the following reaction:

$$\_\_NH_4NO_3 + \_\_Na_3PO_4 \rightarrow \_\_(NH_4)_3PO_4 + \_\_NaNO_3$$

a. If 50 grams of ammonium nitrate and 30 grams of sodium phosphate are available, which is the limiting reagent?

b. How much ammonium phosphate can be produced?

8. Consider the following reaction:

$$3\ CaCO_3 + 2\ FePO_4 \rightarrow Ca_3(PO_4)_2 + Fe_2(CO_3)_3$$

a. How many grams of iron (III) carbonate can be produce from 100 grams of calcium carbonate and 100 g of iron (III) phosphate?

b. What was the limiting reagent?

9. A double replacement reaction occurs between copper (II) sulfate and iron (III) nitrate:

$$\_\_CuSO_4 + \_\_Fe(NO_3)_3 \rightarrow \_\_Fe_2(SO_4)_3 + \_\_Cu(NO_3)_2$$

a. If you place 0.092 mol of iron (III) nitrate in a solution containing 0.158 mol of $CuSO_4$. What is the limiting reactant?

b. How many moles of $Cu(NO_3)_2$ will be formed?

10. In the reaction barium carbonate reacts with nitric acid to form barium nitrate, carbon dioxide and water.

$$\_\_Ba(CO_3) + \_\_HNO_3 \rightarrow \_\_Ba(NO_3)_2 + \_\_CO_2 + \_\_H_2O$$

a. What mass of barium nitrate can be formed by combining 55 g of barium carbonate with 26 g of nitric acid.

b. How many <u>liters</u> of $CO_{2\,(g)}$ will be produced if 15 moles of barium carbonate reacts with 100 grams of nitric acid?

Percent Yield Worksheet

W 325

1) Write a balanced equation for the reaction of tin (IV) phosphate with sodium carbonate to make tin (IV) carbonate and sodium phosphate.

2) If 36 grams of tin (IV) phosphate is mixed with an excess of sodium carbonate, how many grams of tin (IV) carbonate will form?

3) If 29.8 grams of tin (IV) carbonate are actually formed when this reaction goes to completion, what is the percent yield?

4) If 7.3 grams of sodium carbonate are used in the reaction and the result a 74.0% yield, how many grams of sodium phosphate will be formed?

Worksheet: Percent Yield Name________________

1. Chlorobenzene, C_6H_5Cl, is used in the production of chemicals such as aspirin and dyes. One way that chlorobenzene is prepared is by reacting benzene, C_6H_6, with chlorine gas according to the following BALANCED equation.

$$C_6H_6\ (l) + Cl_2\ (g) \rightarrow C_6H_5Cl\ (s) + HCl\ (g)$$

   a. What is the theoretical yield if 45.6 g of benzene react?

   b. If the actual yield is 63.7 g of chlorobenzene, calculate the percent yield.

2. When carbon disulfide burns in the presence of oxygen, sulfur dioxide and carbon dioxide are produced according to the following equation.

$$CS_2\ (l) + 3\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ SO_2\ (g)$$

   a. What is the percent yield of sulfur dioxide if the burning of 25.0 g of carbon disulfide produces 40.5 g of sulfur dioxide?

   b. What is the percent yield of carbon dioxide if 2.5 mol of oxygen react and 32.4 g of carbon dioxide are produced?

PERCENT YIELD PROBLEMS

(http://hhs-chemistry.homewoodcityschools.wikispaces.net/file/view/Percent+Yield+Problems.doc)

1. If, in the reaction below 32 grams of C_2H_6 produces 44 grams of CO_2, what is the % yield?

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

2. If, in the reaction below, 80 grams of Cl_2 produces 38 grams of CCl_4 what is the % yield?

$$CS_2 + 3Cl_2 \rightarrow CCl_4 + S_2Cl_2$$

3. If, in the reaction below, 49 grams of Fe_3O_4 produces a 78.25 % yield of Fe. How many grams are produced?

$$Fe_3O_4 + 4H_2 \rightarrow 3Fe + 4H_2O$$

4. If, in the reaction below, 4 grams of H_2O produces 0.67 grams of HF what is the % yield?

$$CH_3COF + H_2O \rightarrow CH_3COOH + HF$$

5. If, in the reaction below, 28 grams of H_2O produces a 79.59 % yield of HF. How many grams are produced?

$$CH_3COF + H_2O \rightarrow CH_3COOH + HF$$

H-43

Name: ______ ______ Date: ______ Period: ______

Limiting Reagent & Percent Yield Practice Worksheet

1. **When copper (II) chloride reacts with sodium nitrate, copper (II) nitrate and sodium chloride are formed.**
   a. Write the balanced equation for the reaction given above:

   $$CuCl_2 \quad + \quad NaNO_3 \quad \rightarrow \quad Cu(NO_3)_2 \quad + \quad NaCl$$

   b. If 15 grams of copper (II) chloride react with 20 grams of sodium nitrate, how much sodium chloride can be formed?

   c. What is the limiting reagent for the reaction in #2?

   d. How many grams of copper (II) nitrate is formed?

   e. How much of the excess reagent is left over in this reaction?

   f. If 11.3 grams of sodium chloride are formed in the reaction described in problem #2, what is the percent yield of this reaction?

   $$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \text{ percent}$$

2. **When lead (II) nitrate reacts with sodium iodide, sodium nitrate and lead (II) iodide are formed.**
   a. Balance the following equation:

   $$Pb(NO_3)_{2\,(aq)} + \quad NaI_{(aq)} \rightarrow \quad PbI_{2\,(s)} + \quad NaNO_{3\,(aq)}$$

   b. If I start with 25.0 grams of lead (II) nitrate and 15.0 grams of sodium iodide, how many grams of sodium nitrate can be formed?

c. What is the limiting reagent in the reaction described in problem 2?

d. How many grams of lead (II) iodide is formed?

e. How much of the non-limiting reagent will be left over from the reaction in problem #2?

f. If 6 grams of sodium nitrate are formed in the reaction described in problem #2, what is the percent yield of this reaction?

3. 1000 grams of sodium chloride is combined with 2000 grams of barium phosphate.

a. Balance the following equation:

$$6\ NaCl + Ba_3(PO_4)_2 \rightarrow Na_3PO_4 + BaCl_2$$

b. What is the limiting reactant?

c. How many grams of excess reactant are left?

4. A chemist burns 160.0 g of Al in excess air to produce aluminum oxide, Al_2O_3. She produces 260.0 g of solid aluminum oxide.

a. Write a balanced equation for the reaction.

$$Al + O_2 \rightarrow Al_2O_3$$

b. Determine the theoretical yield of Al_2O_3.

c. Determine the percent yield.

Limiting Reactants

A. What is a Limiting Reactant (Limiting Reagent)?

B. Sample Exercise 3.15 – Part of the SO_2 that is introduced into the atmosphere ends up being converted to sulfuric acid. The net reaction is :

$$2\ SO_2\ (g) + O_2\ (g) + 2H_2O\ (l) \rightarrow 2\ H_2SO_4\ (aq)$$

How much sulfuric acid can be formed from 5.0 mol of SO_2, 1.0 mol of O_2, and an unlimited quantity of H_2O?

C. Sample Exercise 3.16 – Consider the following reaction:

$$2\ Na_3PO_4\ (aq) + 3\ Ba(NO_3)_2\ (aq) \rightarrow Ba_3(PO_4)_2\ (s) + 6\ NaNO_3\ (aq)$$

Suppose that a solution containing 3.50 g of Na_3PO_4 is mixed with a solution containing 6.40 g of $Ba(NO_3)_2$. How many grams of $Ba_3(PO_4)_2$ can be formed?

D. What is a <u>theoretical yield</u>?

1. What is the <u>percent yield</u>?

E. Sample Exercise 3.17 – Adipic acid, $H_2C_6H_8O_4$, is a raw material used for the production of nylon. It is made commercially by a controlled reaction between cyclohexane, C_6H_{12}, and O_2:

$$2\ C_6H_{12} + 5\ O_2 \rightarrow 2\ H_2C_6H_8O_4 + 2\ H_2O$$

(a) Assume that you carry out this reaction starting with 25.0 g of cyclohexane, and the cyclohexane is the limiting reactant. What is the theoretical yield of adipic acid?

(b) If you obtain 33.5 g of adipic acid for your reaction, what is the percent yield of adipic acid?

H-47

Name ______________________________ Date __________

Percent Yield / Limiting Reagent Calculations

1. Calculate the percent yield for the reaction:

$$\_\_P_{4\,(s)} + \_\_Cl_{2\,(g)} \rightarrow \_\_PCl_{3\,(l)}$$

   a. If 75.0 g of phosphorus reacts with excess chlorine gas to produce 111.0 g of phosphorus trichloride. What is the percent yield?

2. Calculate the percent yield for an experiment in which 5.50 g of $SOCl_2$ was obtained in a reaction of 5.80 g of SO_2 with 150 g of PCl_5. Use the following equation:

$$\_\_SO_{2\,(l)} + \_\_PCl_{5\,(l)} \rightarrow \_\_SOCl_{2\,(l)} + \_\_POCl_{3\,(l)}$$

3. Consider the reaction:

$$\_\_C_6H_6 + \_\_Br_2 \rightarrow \_\_C_6H_5Br + \_\_HBr$$

   a. What is the theoretical yield of C_6H_5Br if 42.1 g of C_6H_6 react with 73.0 g of Br_2?

   b. If the actual yield of C_6H_5Br is 63.6 g, what is the percent yield?

4. Use the following reaction:

$$\_\_C_4H_9OH + \_\_NaBr + \_\_H_2SO_4 \rightarrow \_\_C_4H_9Br + \_\_NaHSO_4 + \_\_H_2O$$

   a. If 25.0 g of C_4H_9OH react with 35 g of NaBr and 50 g of H_2SO_4 to yield 17.1 g of C_4H_9Br, what is the percent yield of this reaction?

5. Silicon nitride (Si_3N_4) is made by a combining Si and nitrogen gas (N_2) at a high temperature. How many liters of Si gas is needed to react with an excess of nitrogen gas to prepare 125 g of silicon nitride if the percent yield of the reaction is 95.0%?

$$\_\_Si_{(g)} + \_\_N_{2\,(g)} \rightarrow \_\_Si_3N_{4\,(g)}$$

6. Souring of wine occurs when ethanol is converted to acetic acid by oxygen by the following reaction:

$$\_\_C_2H_5OH + \_\_O_2 \rightarrow \_\_CH_3COOH + \_\_H_2O$$

A 1.00 L bottle of wine, labeled as 8.5% ethanol, is found to have a defective seal. Analysis of 1.00 mL showed that there were 0.0274 grams of acetic acid in that 1.00 mL. The density of ethanol is 0.816 g/mL and the density of water is 1.00 g/mL.

   a. What mass of oxygen must have leaked into the bottle?

   b. What is the percent yield for the conversion of ethanol to acetic acid if O_2 is in excess?

7. For the following reaction occurs:

$$P_4 + O_2 \rightarrow P_4O_6$$

A reaction container holds 5.77 g of P_4 and 5.77 g of O_2. If enough oxygen is available then the P_4O_6 reacts further:

$$P_4O_6 + O_2 \rightarrow P_4O_{10}$$

a. What is the limiting reagent for the formation of P_4O_{10}?

b. What mass of P_4O_{10} is produced?

c. What mass of excess reactant is left in the reaction container?

8. Given the following reaction:

$$Mg + HNO_3 \rightarrow Mg(NO_3)_2 + H_2$$

Type of reaction: ______________________________

a. If this reaction starts with 40 grams of magnesium and 50 grams of nitric acid, how many grams of hydrogen gas will be produced?

b. If only 1.7 grams of hydrogen is actually produced, what was my percent yield of hydrogen?

9. Given the following reaction:

$$NaHCO_3 \rightarrow NaOH + CO_2$$

Type of reaction: ______________________________

a. If 25 grams of carbon dioxide gas is produced in this reaction, how many grams of sodium hydroxide should be produced?

b. If 50 grams of sodium hydroxide are actually produced, what was my percent yield?

10. Nickel replaces silver from silver nitrate the following reaction:

$$Ni + Ag(NO_3) \rightarrow Ag + Ni(NO_3)_2$$

a. If you have 22.9 g of Ni and 112 g of $AgNO_3$, which reactant is in excess?

b. What mass of nickel (II) nitrate would be produced given the quantities above?

11. Carbon disulfide is an important industrial substance. Its fumes can burn explosively in air (oxygen) to form sulfur dioxide and carbon dioxide.

$$CS_2 + O_2 \rightarrow CO_2 + SO_2$$

a. If 1.60 mol of CS2 burns with 5.60 mol of oxygen gas, how many moles of the excess reactant will still be present when the reaction is over?

H-51

1. Bromine replaces iodine in magnesium iodide in the following reaction:

$$Br_2 + MgI_2 \rightarrow I_2 + MgBr_2$$

   a. Which is the excess reactant when 560 g of magnesium iodide and 360 g of bromine gas reacts, and what mass remains?

   b. What mass of iodine is formed in the same process?

   c. *Assume the iodine was vaporized at 150 C and 150 kPa, how many liters would it occupy?

THE Bonding Worksheet

THE Bonding Worksheet.doc

Name ______________

Date ______________________

Period____

Complete the table for each of the molecules. * When determining the 3D structure for a molecule using VSEPR or Hybridization consider the underlined atom for the 3D structure also use the underlined atom for determining the number of σ and π bonds.

Molecular Formula	Electron (Lewis) Dot Structure	3D Molecular Shape (**VSEPR label**) *	Hybrid Orbital * (xy^3z^2 designation)	3D Hybridization structure (molecular shape)*	Polar or Nonploar molecule?	# σ bonds*	# π bonds*
CH_4							
BF_3							
XeF_4							

Complete the table for each of the molecules. * When determining the 3D structure for a molecule using VSEPR or Hybridization consider the underlined atom for the 3D structure also use the underlined atom for determining the number of σ and π bonds.

Molecular Formula	Electron (Lewis) Dot Structure	3D Molecular Shape **(VSEPR label)** *	Hybrid Orbital * (xy^3z^2 designation)	3D Hybridization structure (molecular shape)*	Polar or Nonploar molecule?	# σ bonds*	# π bonds*
NH_3							
H_2O							
PF_5							
SF_4							

Name: ______ period: ______

VSEPR Worksheet

1) What is the main idea behind VSEPR theory?

2) For each of the following compounds, determine the bond angles, molecular shapes, and hybridizations for all atoms:

a) carbon tetrachloride

b) BH_3

c) silicon disulfide

d) C_2H_2

e) PF_3

Name: ________ per: ______

sigma (σ) and pi (π) bonds

1) Describe the formation of a sigma (σ) bond.

2) Describe the formation of a pi (π) bond.

3) Complete the following table:

Type of overlap	Type of bond formed
s and s head on	
s and p head on	
p and p head on	
p and p sideways	

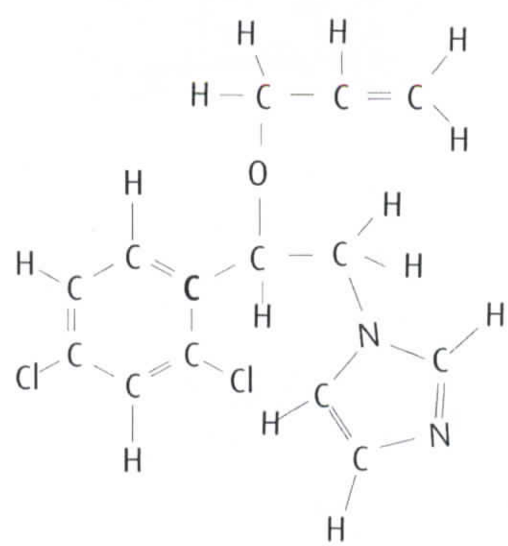
4) Determine the number of sigma (σ) and pi (π) bonds in the following molecules:

a) Cl_2

b) O_2

c) N_2

5) Determine the number of sigma (σ) and pi (π) bond in a molecule of imazalil (pictured below).

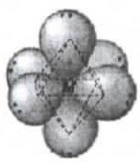


Name ________ per: ________

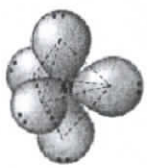
Worksheet 15 - Molecular Shapes

The shapes of molecules can be predicted from their **Lewis structures** by using the **VSEPR** (**V**alence **S**hell **E**lectron **P**air **R**epulsion) model, which states that electron pairs around a central atoms will assume a geometry that keeps them as far apart from each other as possible.

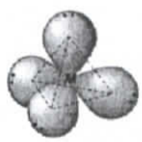
This is illustrated by the drawings below.



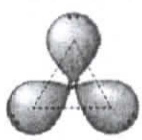
Six groups surrounding a central atom will form an **octahedron**. All of the groups in this structure are at **90°** or **180°** to each other. All positions are equivalent



Five groups will form a **trigonal bipyramid**. The two positions pointing up and down are called the **axial** positions. They are at **180°** to each other, and at **90°** to the other three, **equatorial** positions. The three **equatorial** positions are at **120°** to each other. There is more room in the equatorial positions, and large groups will occupy these positions.



Four groups will form a **tetrahedron**. All of the angles in a tetrahedron are **109.5°**, and all positions are equivalent.



Three groups will form a flat triangle (**trigonal planar**). Each of the angles is **120°** and all positions are equivalent.

Two groups form a straight line (**linear**) with **180°** between them.

How does this apply to Chemistry?

The groups occupying these geometric positions will be either **atoms** bonded to the central atom, or **lone pair electrons** on the central atom.

Lone pair electrons occupy **more** space than bonded electrons, so they will take the **equatorial** position in the **trigonal bipyramid**.

Lone pair electrons will also occupy positions that put them as far apart from each other as possible.

1. Draw the Lewis structure for water, H_2O.

   a) How many "groups" (atoms and lone pairs) surround the central oxygen?

   b) What is the **geometry** of this molecule (look at atoms and lone pairs)? Draw this VSEPR structure next to the Lewis structure.

   c) What is the **shape** of this molecule (look only at the atoms)?

   d) What is the H-O-H bond angle?

   e) Place the partial positive and negative charges on the H and O atoms, based on their relative electronegativities. Is water a **polar** compound?

2. Draw the Lewis structure for NO_2^-.

   a) How many "groups" (atoms and lone pairs) surround the central nitrogen?

   b) What is the **geometry** of this molecule (look at atoms and lone pairs)? Draw this VSEPR structure next to the Lewis structure.

   c) What is the **shape** of this molecule (look only at the atoms)?

   d) What is the O-N-O bond angle?

   e) Place the partial positive and negative charges on the N and O atoms, based on their relative electronegativities. Is NO_2^- a **polar** compound?

Now fill in the missing information in the chart using the structures you have drawn in problems 1 - 3.

compound	**atoms** on central atom	**lone pairs** on central atom	**geometry**	**shape**	**polar**
SF_6				octahedral	
	5	1			
	4		octahedral		
$XeCl_3^-$					
	5	0			
	4	1		seesaw	
BrF_3					
			trigonal bipyramidal	linear	
	4	0			
NH_3					
	2	2		V-shaped (bent)	yes
			trigonal planar		no
	2	1			
CO_2					

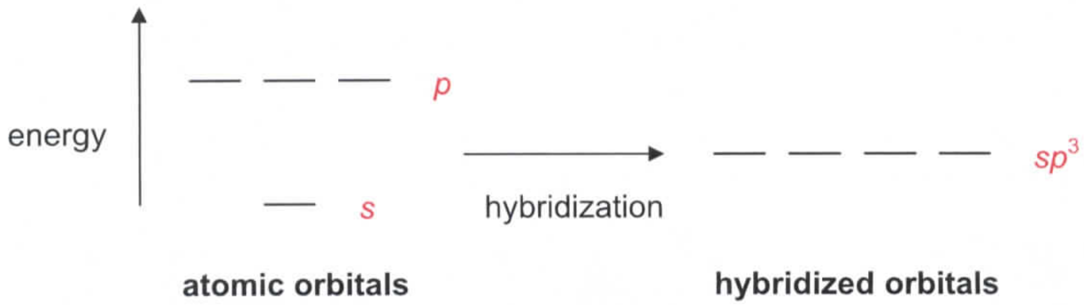
Name: H-59 ____________ Per: ______

Worksheet 14 - Hybridization

When atoms bond to form molecules, they use **molecular orbitals**. These are formed through the **hybridization** of the **atomic orbitals** that we have already discussed, ***s***, ***p***, and ***d*** orbitals.

The **hybridized molecular orbitals** have different shapes and energy levels than the **atomic orbitals**. The number of molecular orbitals created by hybridization depends on the number of atomic orbitals that are mixed to form them.

In forming $\boldsymbol{sp^3}$ hybridized orbitals, **four** atomic orbitals are mixed, one *s* and three *p*. The energy diagram for this process is shown below. The hybridized orbitals are higher in energy than the *s* orbital, but lower in energy than the *p* orbitals.



Carbon has 4 valence electrons. Add these electrons to the atomic and molecular orbitals. This hybridization gives **tetrahedral geometry.**

With this hybridization, C will form **four equivalent σ bonds**.

Draw a similar energy diagram for $\boldsymbol{sp^3}$ hybridized **oxygen**.

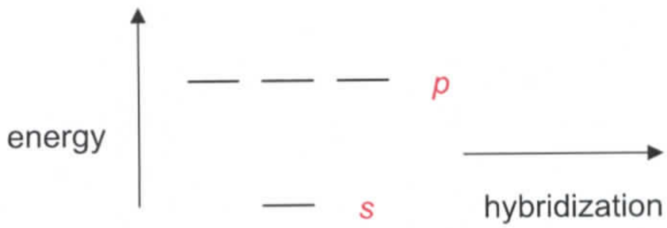
How many σ bonds will be formed?

How are the other sp^3 orbitals used?

Do the same for $\boldsymbol{sp^3}$ hybridized **nitrogen.**

In some Lewis structures, there are only **three** equivalent bonds formed. To create three equivalent hybridized orbitals, mix **three** atomic orbitals.

Draw and name the orbitals formed in this hybridization, then add the electrons for **sulfur**. Since the hybridized orbitals are close in energy, every orbital is filled with one electron before electrons are paired.



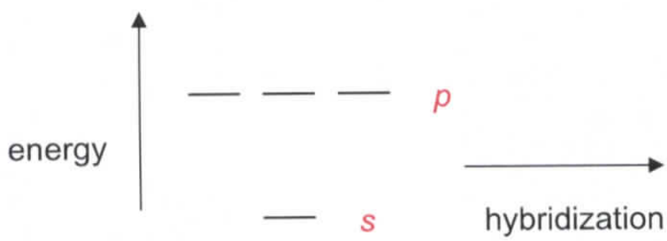
The hybridized orbitals will form _____ σ bond(s).

The unhybridized orbital will form _____ π bond(s).

There will be _____ lone pair(s).

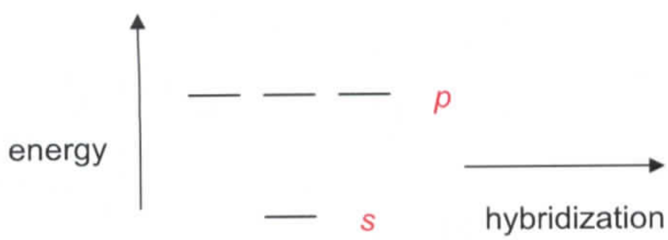
This hybridization gives **trigonal planar** geometry.

In **linear** molecules, like CO_2, the central atom has only **two** equivalent bonding orbitals. Draw the energy levels and name the orbitals formed in this hybridization.



Fill in the electrons for carbon and determine the number and typed of bonds formed.

In CO_2, determine the hybridization of the **oxygen** atoms. Complete the energy diagram for the oxygens. Draw the structure of CO_2.



In atoms with n=3 or larger, the *d* orbitals can also be hybridized. In molecules with **five** molecular orbitals, **five** atomic orbitals are mixed:

___ + ___ + ___ + ___ + ___

This will give **trigonal bipyramidal** geometry and is called ***dsp***^{3} hybridization.

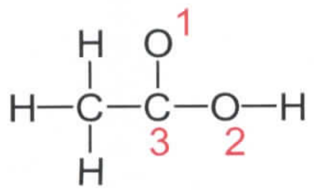
Finally, molecules with **octahedral** geometry, will have ____ molecular orbitals. This hybridization is called ________ .

Shown below is a portion of the chart from **Worksheet 13**. Fill in the **hybridization** for each of the compounds.

compound	bonds	lone pairs	geometry	shape	hybridization
SF_6	6	0	octahedral	octahedral	
NH_3	3	1	tetrahedral	trigonal pyramidal	
ICl_4^-	4	2	octahedral	square planar	
CF_4	4	0	tetrahedral	tetrahedral	
SO_3	3	0	trigonal planar	trigonal planar	
SF_4	4	1	trigonal bipyramidal	seesaw	
CO_2	2	0	linear	linear	
H_2O	2	2	tetrahedral	V-shaped	
NO_2^-	2	1	trigonal planar	V-shaped	

Fill in the chart below and then complete the Lewis structures for the molecules shown below and fill in those charts.

element	Lewis symbol	# bonds	# lone pairs
C			
N			
H			
O			
Halogen			

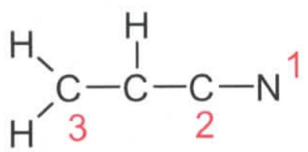


acetic acid

σ bonds ____

π bonds _____

atom #	bond angle	hybridization
1		
2		
3		

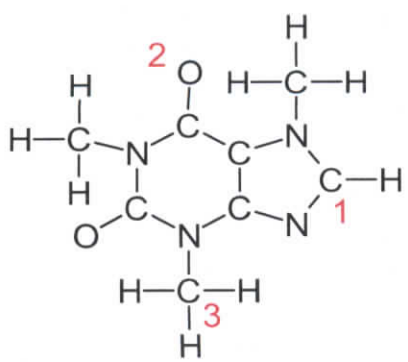


acrylonitrile

σ bonds ____

π bonds _____

atom #	bond angle	hybridization
1		
2		
3		

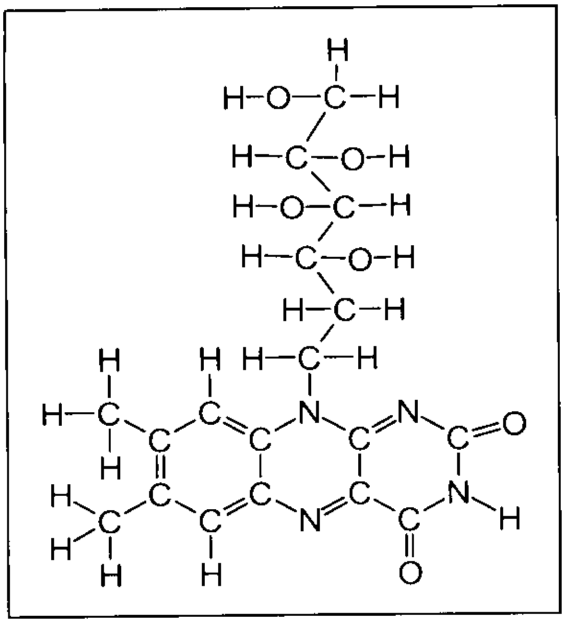


caffeine

σ bonds ____

π bonds _____

atom #	bond angle	hybridization
1		
2		
3		



The molecule shown to the left is riboflavin (vitamin B2). Answer the following questions about its structure.

a) how many carbons are sp^3 hybridized?

sp^2 hybridized?

sp hybridized?

b) How many nitrogens are sp^3 hybridized?

sp^2 hybridized?

sp hybridized?

c) How many oxygens are sp^3 hybridized?

sp^2 hybridized?

sp hybridized?

d) How many σ bonds are there in total?

e) How many π bonds are there in total?

f) Which of the three rings are **planar**?

The acetate ion, $C_2H_3O_2^-$, has both oxygens bonded to the same carbon.

a) Draw the Lewis structure and all resonance forms.

b) Label the hybridization around each carbon.

c) Pick one resonance structure and label the hybridization of each oxygen.

d) How many σ and π bonds are present?

e) Which atom carries the formal negative charge?

Formal Charge

- Formal charges are used when there is more than one possible Lewis structure for a molecule.
- The Formal charge of an atom equals the number of valence electrons in the isolated atom, minus the number of electrons assigned to the atom in the Lewis Structure
- Most stable lewis structure is one where all charges equal zero or the negative charge resides on the most electronegative element.
- Rules:
    - Valence e- for that atom minus left over electrons around atom and ½ e- in the bond

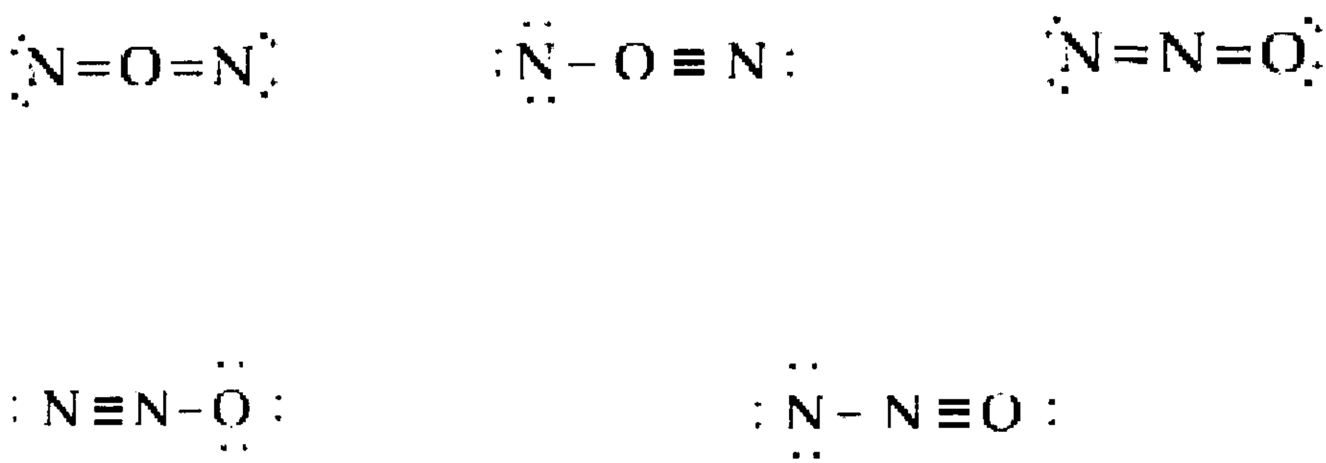
For example:

Valence e-

\- e- assigned

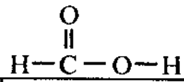
Formal Charge

Example #2; Use the same process as above.



HYBRIDIZATION REVIEW WORKSHEET

1. What is the number of sigma (σ) and pi (π) bonds and the hybridization of the carbon atom in



	Sigma	Pi	Hybridization
A.	4	1	sp^2
B.	4	1	sp^3
C.	3	2	sp^3
D.	3	1	sp^2

2. What is the best description of the carbon-oxygen bond lengths in CO_3^{2-}?

   A. One short and two long bonds

   B. One long and two short bonds

   C. Three bonds of the same length

   D. Three bonds of different lengths

3. Which allotropes contain carbon atoms with sp^2 hybridization?

   I. Diamond

   II. Graphite

   III. C_{60} fullerene

   A. I and II only

   B. I and III only

   C. II and III only

   D. I, II and III

4. Which of the following species is (are) planar (has (have) all the atoms in one plane)?

   I. CO_3^{2-}

   II. NO_3^-

   III. SO_3^{2-}

   A. I only

   B. II only

   C. I and II only

   D. II and III only

10. Which statement is correct about multiple bonding between carbon atoms?

A. Double bonds are formed by two π bonds.

B. Double bonds are weaker than single bonds.

C. π bonds are formed by overlap between s orbitals.

D. π bonds are weaker than sigma bonds.

Short Answers:

1. (i) List the following substances in order of increasing boiling point (lowest first).

CH_3CHO C_2H_6 CH_3COOH C_2H_5OH

.... **(2)**

(ii) State whether each compound is polar or non-polar, and explain the order of boiling points in (c)(i).

....

....

....

....

....

....

....

.... **(8)**

2. Atomic orbitals can mix by hybridization to form new orbitals for bonding.

Identify the type of hybridization present in each of the **three** following molecules.
Deduce and explain their shapes.

(i) OF_2

....

....

....

....

....

.... **(3)**

4. (i) Explain the meaning of the term *hybridization*.

....

....

.... **(1)**

(ii) Discuss the bonding in the molecule CH_3CHCH_2 with reference to

- the formation of σ and π bonds
- the length and strength of the carbon-carbon bonds
- the types of hybridization shown by the carbon atoms

....

....

....

....

....

....

....

....

(6)

5. The boiling points of the hydrides of group 6 elements increase in the order

$$H_2S < H_2Se < H_2Te < H_2O.$$

Explain the trend in the boiling points in terms of bonding.

....

....

.... **(6)**

Name_________________________________ Block____

Honors Chemistry

Remember that this is a MOLE fraction so if you are given mass **YOU MUST FIRST CONVERT TO MOLES.**

For example: A mixture consists of 27 g of Ar and 15 g of O_2. What is the mole fraction of each component?

Before we do anything we need to convert to moles:

$27\text{ g Ar} \times \frac{1\text{ mol}}{40\text{ g}} = 0.675\text{ mol Ar}$ $\quad$ $15\text{ g }O_2 \times \frac{1\text{ mol}}{32\text{ g}} = 0.469\text{ mol }O_2$

1. Find the total moles:

0.675 mol Ar + 0.469 mol O_2 = 1.14 mol total

2. Divide the moles of the desired by the total moles:

$\chi_{Ar} = \frac{0.675\text{ mols Ar}}{1.14\text{ mol total}} = 0.592$ (no units!)

$\chi_{O2} = \frac{0.469\text{ mols }O_2}{1.14\text{ mol total}} = 0.411$ (no units!)

Your turn:

A gaseous mixture made from 10 g of nitrogen and 5 g of chlorine. What is the mole fraction of each component?

Given:

Convert to moles!

1. Find the total moles:

2. Divide the moles of the desired by the total moles:

Name________________________________ Block____

Honors Chemistry

NOW: Use the mole fraction and total pressure to find partial pressure.
A gas mixture is created by blending 0.2 mols CO_2, 2.0 mols O_2 and 7.8 mols N_2. If the total pressure is 750 mmHg, calculate the partial pressure of oxygen.

Given: 0.2 mols CO_2, 2.0 mols O_2 , 7.8 mols N_2

750 mm Hg = P_t

Equations: $\chi = \frac{\text{moles of desired component}}{\text{total moles of mixture}}$

$\chi \cdot P_t = P$

1. Find χ_{O2}

Total moles = 0.2 mols CO_2 + 2.0 mols O_2 + 7.8 mols N_2 = 10 moles total

$\chi_{O2} = \frac{2.0 \text{ mol}O_2}{10 \text{ moles total}} = 0.2$

2. Find P_{O2}

$\chi \cdot P_t = P \rightarrow$ 0.2 x 750mmHg = 150 mmHg

Your turn:
A gas mixture is created by blending 5.2 mols CO_2, 13.1 mols O_2 and 2.4 mols N_2. If the total pressure is 710 mmHg, calculate the partial pressure of Oxygen

Given: 5.2 mols CO_2, 13.1 mols O_2 , 2.4 mols N_2

710 mm Hg = P_t

Want: P_{O2}

Equations: $\chi = \frac{\text{moles of desired component}}{\text{total moles of mixture}}$

$\chi \cdot P_t = P$

1. Find χ_{O2}

Total moles =

χ_{O2} =

2. Find P_{O2}

$\chi \cdot P_t = P \rightarrow$

Remember if given grams first convert to moles and complete the calculations!!

Name______________________________ Block____

Intro to Mole Fractions and Partial Pressure

You didn't think you were going to get away with using Dalton's law of partial pressures alone did you??

Dalton's Law of partial pressures: $P_t = P_1 + P_2 + P_3 + ...P_n$

How can you calculate the partial pressure of a gas using P_t?

$$\chi \cdot P_t = P$$

χ - mole fraction
P_t – total pressure
P – partial pressure

A mole fraction, χ, is a way of expressing the components of a mixture with moles.

$$\chi = \frac{\textbf{moles of desired component}}{\textbf{total moles of mixture}}$$

For example, given this problem:
A gas mixture is composed of 3 moles O_2 and 7 mols N_2. Find the mole fraction of O_2.
Given: 3 mols O_2
7 mols N_2
Want : χ_{O2}
Equation: $\chi = \frac{\text{moles of desired component}}{\text{total moles of mixture}}$

1. Find total moles in the mixture: 3 mol O_2 + 7 mol N_2 = 10 mol of mixture

2. Divide the moles of the desired by the total moles:

$$\chi_{O2} = \frac{3 \text{ mols } O_2}{10 \text{ total moles}} = 0.3 \text{ (no units!)}$$

Your turn:
A mixture has 4 mols of CO_2, 5 mols of N_2 and 12 mols of He. What is the mole fraction of CO_2? What is the mol fraction of He?

Given:

Want:
Equation:

1. Find the total moles:

2. Divide the moles of the desired by the total moles:

Worksheet –Dalton's Law of Partial Pressure

Name______________

Period______

Date_______

Directions: Show all work including the formula used. Box in your answers

Dalton's Law

1. What is the pressure of a mixture of nitrogen (N_2) and oxygen (O_2) if partial pressure of N_2 is 594 mm Hg and the partial pressure of O_2 is 165 mm Hg?

2. What is the partial pressure of hydrogen gas in a mixture of hydrogen and helium if the total pressure is 600 mm Hg and the partial pressure of helium is 439 mm Hg?

3. Find the partial pressure of carbon dioxide in a gas mixture with a total pressure of 30.4 kPa if the partial pressures of the other two gases in the mixture are 16.5 kPa and 3.7 kPa.

4. What is the partial pressure of water vapor in an air sample when the total pressure is 1.00 atm, the partial pressure of nitrogen is 0.79 atm, the partial pressure of oxygen is 0.20 atm, and the partial pressure of all other gases in the air is 0.0044 atm?

5. The pressure of a gas is 12.9 mm Hg. Express this value in each of the following units.
   a. torr
   b. atmosphere
   c. kilopascal

6. At an ocean depth of 250 feet, the pressure is about 8.4 atm. Convert the pressure to mm Hg and kPa units.

7. The atmospheric pressure in Denver, Colorado, is usually about 84.0 kPa. What is this pressure in atm and torr units?

8. What is the total pressure in atmospheres of a mixture of three gases with partial pressures of 0.118 atm, 35.6 kPa, and 167 mm Hg?

9. What is the partial pressure of N_2 if the total pressure of a container with both nitrogen and carbon dioxide gas is 1.604 atm and carbon dioxide's partial pressure is 41.80 kPa?

10. Air is composed of nitrogen, oxygen, argon, and trace gases. If the total atmospheric pressure is 760 torr, and the partial pressures of nitrogen is 592 torr, oxygen is 160 torr, and argon is 7 torr, what is the partial pressure in mm Hg of the trace gases?

H-73

Name______________________________ Block____

Honors Chemistry

Practice Problems:

Complete these problems using the handout and we will review the information in class.

1. A tank contains 5.00 moles of O_2, 3.00 moles of neon, 6.00 moles of H_2S, and 4.00 moles of argon at a total pressure of 1620.0 mm Hg. What is the mole fraction of each gas?
What is the partial pressure of each gas?

2. A mixture of 14.0 grams of hydrogen, 84.0 grams of nitrogen, and 2.0 grams of oxygen are placed in a flask. The total pressure in the flask is 1604 mm Hg. What is the mole fraction of each gas?
What is the partial pressure of each component?

1 atm = 101.3 Kpa = 760 mm Hg = 760 torr

Name: ____________________________

Ideal Gas Law I

PV=nRT

Conversions:

1 atm = 760 torr

= 760 mmHg

= 101.3 kPa

Required Units:

n - Moles   T - Kelvin

V - Liters   P - kPa

R - 8.31 $\frac{kPa \cdot L}{K \cdot mol}$ gas constant

R - .0821 $\frac{atm \cdot L}{K \cdot mol}$

1. How many moles of oxygen will occupy a volume of 2.5 liters at 1.2 atm and 25°C?

2.What volume will 2.0 moles of nitrogen occupy at 720 torr and 20°C?

3.What pressure will be exerted by 25g of CO_2 at 25°C and 500 ml?

4.What temperature will 5g of Cl_2 exert a pressure of 900 torr at a volume of 750ml?

5***What is the density of NH_3 at 800 torr and 25°C?

6. H-75 If the density of a gas is 1.2 g/L at 745 torr and 20°C, what is its molecular mass?

7. How many moles of nitrogen gas will occupy a volume of 347 mL at 6680 torr and 27°C?

8. What volume will 454 grams of hydrogen occupy at 1.05 atm and 25°C?

9. Find the number of grams of CO_2 that exert a pressure of 785 torr at a volume of 32.5 L and a temperature of 32°C?

10. An elemental gas has a mass of 10.3g, if the volume is 58.4 L and the pressure is 758 torr at a temperature of 2.5°C, what is the gas?

Name H-76 ______________________________ Date ______________

PV=nRT II

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

Required Units:
n - Moles
T - Kelvin
V - Liters
P - kPa
R - 8.31 $\frac{kPa \cdot L}{K \cdot mol}$
R - .0821 $\frac{atm \cdot L}{K \cdot mol}$

1. 5.00 L of a gas is known to contain 0.965 mol. If the amount of gas is increased to 1.80 mol, what new volume will result (at an unchanged temperature and pressure)?

2. A cylinder with a movable piston contains 2.00 g of helium, He, at room temperature. More helium was added to the cylinder and the volume was adjusted so that the gas pressure remained the same. How many grams of helium were added to the cylinder if the volume was changed from 2.00 L to 2.70 L? (The temperature was held constant.)

3. What will be the volume of 10.00 grams of sulfur dioxide gas (SO_2) at 35°C and 740 mmHg.

4. A weather balloon is filled with 1000L of hydrogen gas at 732.6 mmHg and 10°C. How many moles of hydrogen molecules are present in the balloon?

5. A 28 gram sample of nitrogen is sealed in a 6.00 L container and heated to a temperature of 327°C. What is the pressure of the gas in Kpa?

6. 22400 Liters of propane gas at 0.95 atm and 15°C must be pressurized and condensed to fill a typical barbeque grill tank. How many grams of propane (C_3H_8) are now inside the container?

7. Determine the density of carbon dioxide in grams/liter at STP. What is the density of carbon dioxide in grams/liter at a temperature of 120°C.

8. How many moles of $CO_{2(g)}$ is in a 5.6 L sample of CO_2 measured at STP?

9. A) Calculate the volume of 4.50 mol of $SO_{2(g)}$ measured at STP. b) What volume would this occupy at 25°C and 150 kPa?

10. How many grams of $Cl_{2(g)}$ can be stored in a 10.0 L container at 1000 kPa and 30°C?

11. At 150°C and 100 kPa, 1.00 L of a compound has a mass of 2.506 g. Calculate its molar mass.

12. 5.98 mL of an unknown gas weighs 0.081 g at STP. Calculate the molar mass of the gas. Can you determine the identity of this unknown gas?

13. How many moles of H_2 is in a 3.1 L sample of H_2 measured at 300 kPa and 20° C?

14. How many grams of O_2 are in a 315 mL container that has a pressure of 12 atmat 25° C?

15. What pressure will be exerted by 20.16g hydrogen gas in a 7.5L cylinder at 20°C?

16. A 50L cylinder is filled with argon gas to a pressure of 10130.0kPa at 30°C. How many moles of argon gas are in the cylinder?

17. To what temperature does a 250mL cylinder containing 0.40g helium gas need to be cooled in order for the pressure to be 253.25kPa?

18. How many moles of gas does it take to occupy 120 liters at a pressure of 2.3 atm and a temperature of 340 K

19. What is the pressure in a 50 L container that holds 45 moles of a gas at 200 C?

20. It is not safe to put aerosol canisters in a campfire because the heat causes the pressure inside the canister to increase and explode. If a 1 L canister that holds 2 moles of gas is placed on a campfire and allowed to heat up to 1400 C. What is the pressure inside the container?

21. How many moles of gas are in a 30 liter scuba canister if the temperature is 300 K and the pressure is 200 atm?

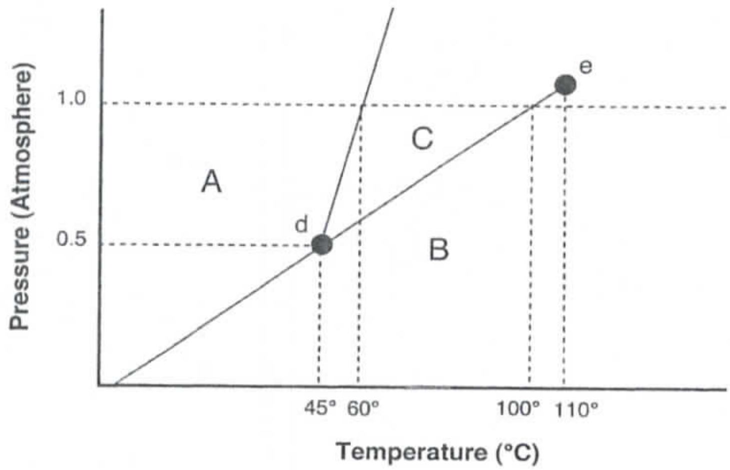
22. A balloon with a maximum volume of 100 L is filled completely with 3 moles of oxygen gas at a pressure of 1 atm, what is the temperature inside the balloon?

23. A sealed piston has a volume of 100 ml at STP. The piston is depressed until volume is .1 ml and the pressure goes up to 1000 atm, what is the new temperature? Assume the number of moles remains constant.

Name ______________________________ Period __________

PHASE DIAGRAM WORKSHEET

Part A – Generic Phase Diagram.
Answer the questions below in relation to the following generic phase diagram.



1. Which section represents the solid phase? ________
2. What section represents the liquid phase? ________
3. What section represents the gas phase? ________
4. What letter represents the triple point? ________

   In your own words, what is the definition of a triple point?

5. What is this substance's normal melting point, at 1 atmosphere of pressure? ________
6. What is this substance's normal boiling point, at 1 atmosphere of pressure? ________
7. Above what temperature is it impossible to liquefy this substance, no matter what the pressure? _____
8. At what temperature and pressure do all three phases coexist? ________________
9. At a constant temperature, what would you do to cause this substance to change from the liquid phase to the solid phase? ______________________________
10. What does sublimation mean? ______________________________

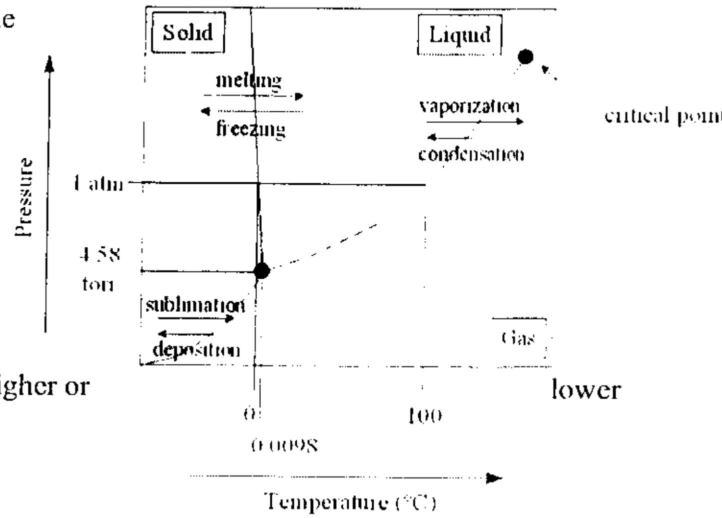
____

Part B – Phase Diagram for Water.

11. At a pressure of 1 atmosphere, what is the normal freezing point of water? ________

12. What is the normal boiling point of water, at one atmosphere of water? ________

13. In Albuquerque, we live approximately 5,500 feet above sea level, which means the normal atmospheric pressure is less than 1 atm. In Albuquerque, will water freeze at a lower temperature or a higher temperature than at 1 atmosphere? _________ Will water boil at a higher or lower temperature, than at 1 atmosphere? ________



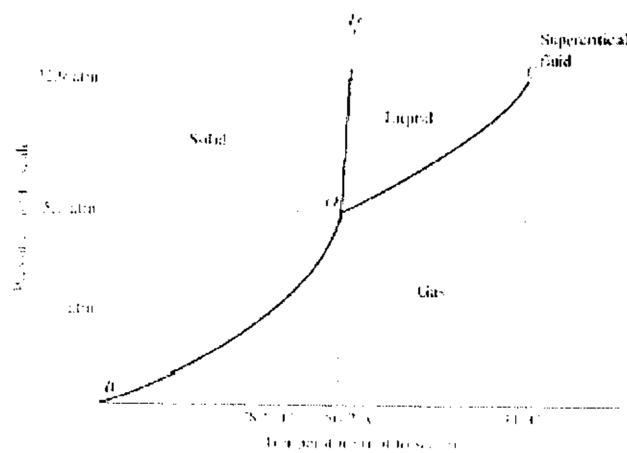
Part C – Phase Diagram for Carbon Dioxide.

14. At 1 atmosphere and room temperature (25°C), would you expect solid carbon dioxide to melt to the liquid phase, or sublime to the gas phase? _________

15. Some industrial processes require carbon dioxide. The carbon dioxide is stored on-site in large tanks as liquid carbon dioxide. Assuming we lived at sea level (1 atm), how could carbon dioxide be liquefied? ____________________

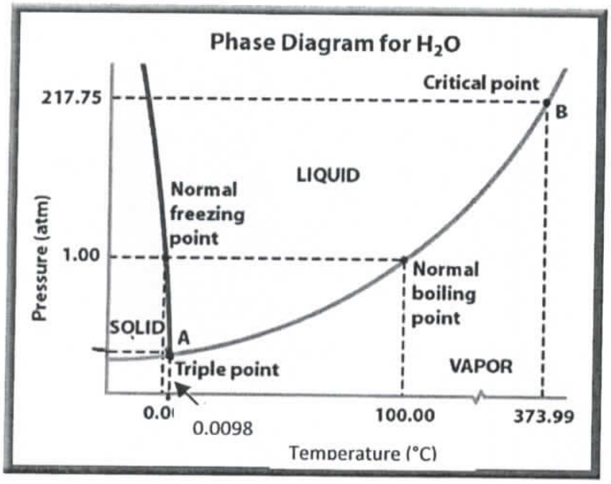
____

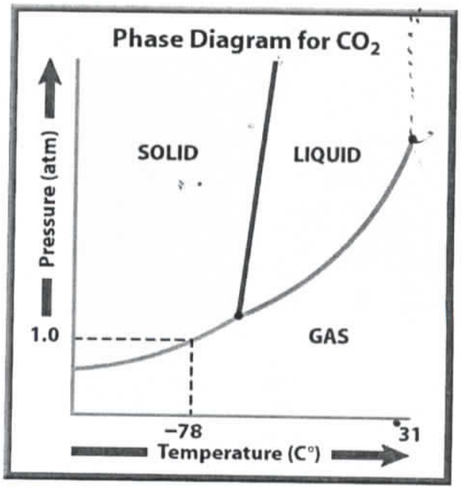
____



PHASE DIAGRAMS

Temperature and ________________________ control the phase of a substance. A phase diagram is a graph of pressure versus temperature that shows in which phase a substance exists under different conditions of temperature and pressure. A phase diagram typically has ______ regions, each representing a different phase and three curves that ____________________________ each phase.





The points on the curves (lines) indicate conditions under which two phases coexist. The critical point indicates the critical pressure and the critical temperature above which a substance cannot exist as a ____________________. The triple point is the point on a phase diagram that represents the temperature and pressure at which three phases of a substance can ______________________________. The ______________________________ slope of the solid-liquid line in the phase diagram for water indicates that the solid floats on its liquid.

Phase Diagram Questions

1. What phase change occurs for CO_2 at –100 °C and 1 atm pressure as it is heated to room temperature?

2. What phase change happens to water at 1 atm as the temperature rises from –15°C to 60°C?

3. What state of matter is water at 50°C and 20 atm?

4. At what temperature does the triple point occur for water?

5. At what temperature does the critical point occur for carbon dioxide?

6. At standard pressure and –78°C, what two phase changes can occur for carbon dioxide?

7. What state of matter is carbon dioxide at –80°C and 2atm ?

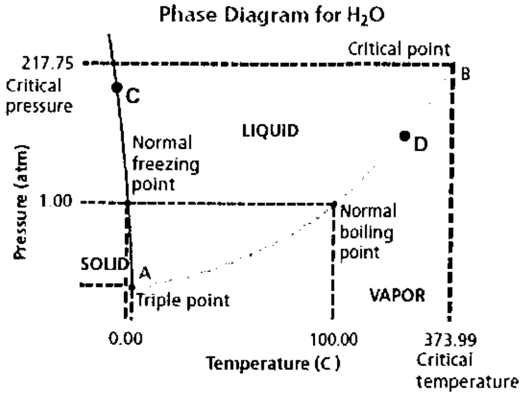
X **PHASE DIAGRAMS WORKSHEET** Name ______________________ Period ____

1. What variables are plotted on a phase diagram?

2. How many phases are represented in a phase diagram? What are they?

3. Use the phase diagram for water to complete the following table.

Temperature (°C)	Pressure (atm)	Phase
200	1	
2	1	
150	100	
−2	0.001	
30	0.8	
0.00 < T < ______	1	liquid
100.00	P < ______	vapor



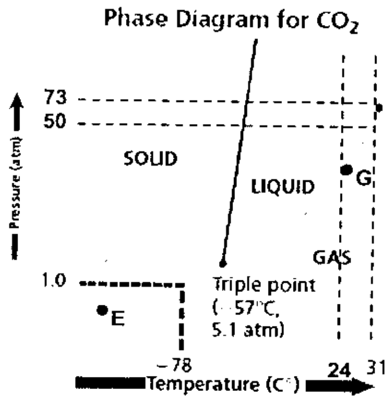
4. What phases of water coexist at point C in water's phase diagram?

5. What two phase changes occur at point D in the phase diagram for water?

6. What is the critical temperature of water?

7. What pressure is at water's normal boiling point?

8. What occurs at the triple point?



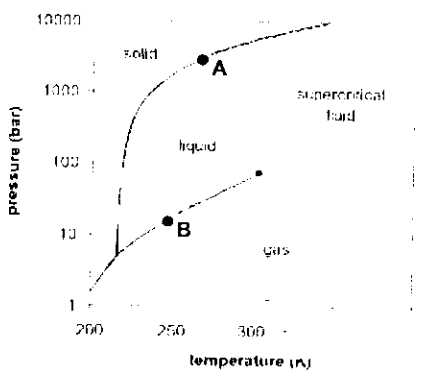
9. Look at the phase diagram for carbon dioxide. Above which pressure and temperature is carbon dioxide unable to exist as a liquid?

10. At which pressure and temperature do the solid, liquid, and gaseous phases of carbon dioxide coexist?

11. What two phase changes occur at point E in the phase diagram for carbon dioxide?

12. What phases of water coexist at point G in CO_2's phase diagram?

13. What phase change occurs as carbon dioxide moves from -78°C to 24°C at a pressure of 50 atm?



Use the phase diagram to the left to answer questions 14-18.

14. What is the temperature at which the triple point occurs?

15. What 2 phase changes occur at Point A?

16. What phase change does the substance at 100 bars undergo as the temperature decreases from 250 K to 200 K?

17. What is the pressure at which the critical point occurs?

18. What 2 phase changes occur at Point B?

$q = mC\Delta T$
$q = mH_f$
$q = mH_v$

$C_{water} = 4.18 \frac{J}{g \cdot K}$ "=" $1.00 \frac{cal}{g \cdot K}$ "

Section 1. Specific Heat

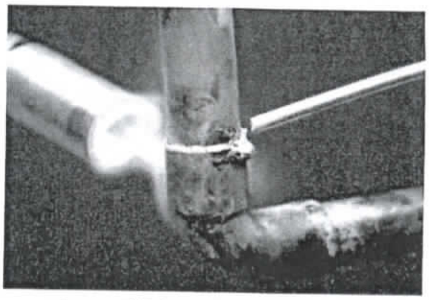
1. A piece of aluminum metal absorbs 455 J of energy, and its temperature rises by 26.0 degrees Celsius. Calculate the mass of this piece of metal. C_{Al} = 0.90 J/(g·°C)

2. A scientist heats a 2.58 gram piece of silver metal, and measures its temperature change from 24 °C to 94 °C. How much energy was absorbed by the silver? C_{Ag} = 0.24 J/(g·°C)

$q = mC\Delta T$

$q = q$
$m = 2.58g$
$C = 0.24 J/g°C$
$\Delta T = T_f - T_I = 94 - 24 = 70°C$

3. Plumbers fuse copper pipe together by the process of *soldering*; two pieces of copper pipe are held together, and heated to a high temperature with a propane torch. At the correct temperature, solder is held to the connection, and it flows into the joint, sealing the two pieces together. Solder acts like a glue that flows into the joint.



It can be very difficult to solder a copper pipe connection, *if there is water still inside the pipes*. Use the idea of *specific heat capacities* to explain why it can be very difficult to reach the required high temperature.

4. Calculate the specific heat of glass from the following data. The temperature of a piece of glass with a mass of 65.0 grams increases by 26.0 °C when it absorbs 840.0 J of energy.

Section 2. Thermochemical Equations

Heat is a product Exo ΔH −
reactant Endo ΔH +

1. Calculate the energy released when 0.25 moles of ethene reacts according to the reaction

$$1\ C_2H_4 + 3\ O_2 \rightarrow 2\ H_2O + 2\ CO_2 + 1401\ kJ$$

2. Is the reaction in question 1 above exothermic or endothermic? Re-write the equation, using ΔH notation.

3. Phosphorus pentachloride absorbs 87.9 kJ of energy per mole when it decomposes into phosphorus trichloride plus chlorine gas. Write a thermochemical equation for this reaction.

4. Calculate the energy absorbed when 3.50 moles of chlorine gas are produced in question 3 above.

5. Ammonia (NH_3) reacts with oxygen in the air, forming nitrogen gas plus water. Each mole of ammonia that reacts, produces 382.5 kJ of energy. Write a thermochemical equation for this reaction.

6. Calculate the energy absorbed when 155.0 grams of aluminum oxide reacts according to the reaction

$$2\ Al_2O_3 \rightarrow 4\ Al + 3\ O_2 \qquad \Delta H = +\ 3352\ kJ$$

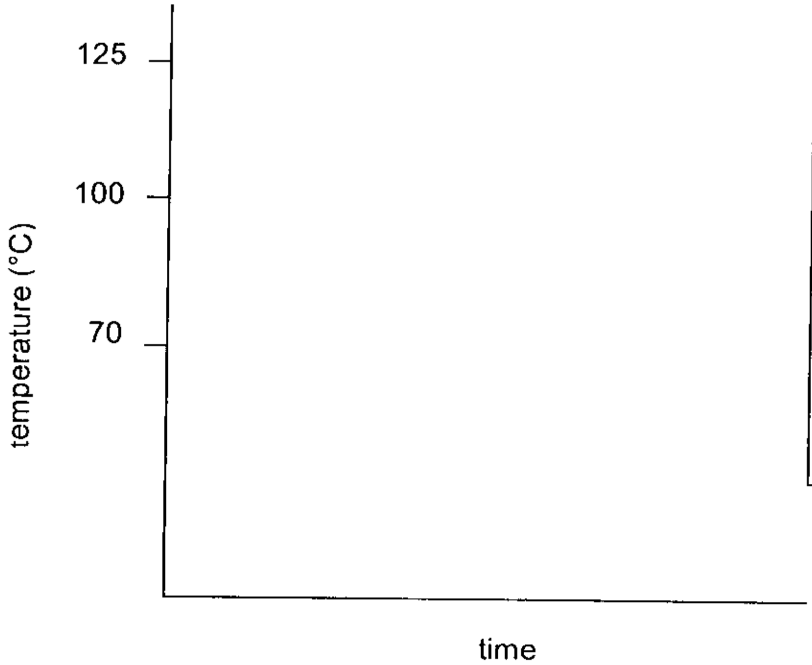
Section 3. Heat in Changes of State

A sample of water (mass = 12.50 g) is heated from 70.0 °C to 125.0 °C.

1. Describe the <u>changes in the state of matter</u> during this heating process.

2. Is this process exothermic or endothermic?

3. Sketch a heating curve of this process.



Specific Heat Capacities and Phase Change Enthalpies:

$C_{liquid\ H2O}$ = 4.18 J/g·°C
$C_{vapor\ H2O}$ = 2.02 J/g·°C
$C_{solid\ H2O}$ = 2.06 J/g·°C
C_{gold} = 0.128 J/g·°C
C_{iron} = 0.448 J/g·°C
$\Delta H_{vap\ H2O}$ = 2260 J/g
$\Delta H_{fus\ H2O}$ = 334 J/g
$\Delta H_{fus\ iron}$ = 14.9 kJ/mol
$\Delta H_{fus\ Pd}$ = 162 J/g

4. (a) Calculate the energy required for the heating of the liquid water.

   (b) Calculate the energy required for the vaporization of the water.

   (c) Calculate the energy required for the heating of the water vapor.

   (d) Sum up the total energy required to heat from 70.0 °C to 125.0 °C.

5. Iron melts at 1535 °C. Calculate the energy that needs to be added to a 5.58 kg sample of iron, to heat it from 995 °C to 1535 °C, and then completely melt the sample.

6. To solidify 0.424 mol Pd, how much heat must be removed?

7. Addition of 8.5 kJ of energy to a boiling water sample will vaporize how many moles H_2O?

8. A piece of gold with a mass of 21.5 g at a temperature of 95.0 °C is dropped into an insulated calorimeter containing 125.0 g of water at 22.0 °C. What will be the final temperature of the system when it reaches equilibrium?

(Repeat quietly to yourself…algebra is my friend, algebra is my friend, algebra is my friend…)

$C_{Au} = 0.128\ J/g\,°C$

HONORS CHEMISTRY
WORKSHEET 8c. Name: ______________________

FINISH AT HOME ANY PROBLEMS THAT WE DO NOT FINISH IN CLASS

Specific heat capacities: **SHOW ALL WORK**

Water	1.00 cal / g °C = 4.18 J / g K	1 cal = 4.18 J
Aluminum	.215 cal / g °C = 0.897 J / g K	
Iron	.107 cal / g °C = 0.449 J / g K	
Gold	.031 cal / g °C = 0.129 J / g K	

1. Calculate the energy required to raise the temperature of 230. g of water from 20.0° C to 35.0° C.

2. A 250.0 g gold statue at 21.0° C is placed in a tank containing 1105. g of water at an initial temperature of 80° C. If no heat enters or leaves the system, find the final equilibrium temperature.

3. Some students heated up a 31.2 g aluminum bar in a beaker of boiling water at 100.0° C.
After it reached equilibrium with the water, they transferred the aluminum to a beaker containing 236.0 g of room temperature water at 21.0° C.
When the aluminum and water reached thermal equilibrium, the final temperature was 23.0° C.
a. Use the students' data to find the specific heat capacity of aluminum.
b. Compare this value with the value at the top of this page and calculate a percentage error.
c. What is the most likely source of experimental error?

4. Use the following data to calculate the energy needed to convert 0.500 kg of ice at -20° C into steam at 250° C.
Specific heat capacity of ice = 2.03 J / g °C = 2.03 J/g K
Specific heat capacity of water = 4.18 J/g K
Specific heat capacity of steam = 2.0 J/g K
Heat of fusion at 0° C = 6.02 kJ/mol
Heat of vaporization at 100° C = 40.7 kJ/mol

Name ______________________________ Date ____________________

smaller - faster
larger - slower

Grahams Law of Diffusion/Effusion I

$$\frac{Rate_A}{Rate_B} = \sqrt{\frac{Molar\ Mass_B}{Molar\ Mass_A}}$$

1. Calculate the rate of effusion of nitrogen gas (N_2) to oxygen gas (O_2). How does this compare to the density of the gases?

2. If equal amounts of helium and argon are placed in a porous container and allowed to escape, which gas will escape faster and how much faster?

3. What is the molecular weight of a gas which diffuses 1/50 as fast as hydrogen?

4. How much faster does hydrogen escape through a porous container than sulfur dioxide?

5. Compare the rate of diffusion of carbon dioxide (CO_2) & ozone (O_3) at the same temperature.

6. *** Two porous containers are filled with hydrogen and neon respectively. Under identical conditions, 2/3 of the hydrogen escapes in 6 hours. How long will it take for half the neon to escape?

7. If the density of hydrogen is 0.090 g/L and its rate of diffusion is 5.93 times that of chlorine, what is the density of chlorine? Use Grahm's Law to determine the molecular mass of chlorine.

8. 2.278×10^{-4} mol of an unidentified gaseous substance effuses through a tiny hole in 95.70 s. Under identical conditions, 1.738×10^{-4} mol of argon gas takes 81.60 s to effuse. What is the molar mass of the unidentified substance?

9. A compound composed of carbon, hydrogen, and chlorine diffuses through a pinhole 0.411 times as fast as neon. Select the correct molecular formula for the compound:

a) $CHCl_3$
b) CH_2Cl_2
c) $C_2H_2Cl_2$
d) C_2H_3Cl

10. Which pair of gases contains one which effuses at twice the rate of the other in the pair?

A. He and Ne
B. Ne and CO_2
C. He and CH_4
D. CO_2 and HCl
E. CH_4 and HCl

11. If a molecule of CH_4 diffuses a distance of 0.530 m from a point source, calculate the distance (in meters) that a molecule of N_2 would diffuse under the same conditions for the same period of time.

12. What is the rate of effusion for a gas that has a molar mass twice that of a gas that effuses at a rate of 3.62 mol/min?

13. Calculate the rate of effusion of NO_2 compared to SO_2 at the same temperature and pressure.

14. Assume you have a sample of hydrogen gas containing H_2, HD, and D_2 that you want to separate into pure components. What are the various ratios of relative rates of effusion? Note: D= Deuterium (heavy hydrogen) a hydrogen atom with a weight of 2 a.m.u. ^2H

15. ***A 3.00 L sample of helium was placed in container fitted with a porous membrane. Half of the helium effused through the membrane in 25 hours. A 3.00 L sample of oxygen was placed in an identical container. How many hours will it take for half of the oxygen to effuse though the membrane?

H-91

Name: ____________________

Honors Chemistry

Colligative Properties Worksheet

$K_f = -1.86\ °C/m$ $K_b = 0.51°C/m$ (these values are for water)

1. What is the molality of a solution of 1.25 moles of sugar dissolved in 0.750 kg of water? What is the boiling point and freezing point of the resulting solution?

2. Calculate the new freezing point of water when 72.0 grams of $BaCl_2$ are added to 175 grams of water.

3. What is the boiling point of a solution made by dissolving 31 g of NaCl in 559 g of water?

4. Calculate the boiling point of an ionic solution containing 29.7 g Na_2SO_4 and 84.4 g water.

5. Calculate the freezing point of a non-ionizing antifreeze solution containing 388g ethylene glycol, $C_2H_6O_2$, and 409 g of water.

6. Calculate the boiling point of a solution made from 227 g of $MgCl_2$ dissolved in 700. g of water.

7. Calculate the molality of a water solution if the freezing point is -9.3°C.

8. Calculate the molality of a water solution if the boiling point is 103.12°C.

9. If 4.18 g of a nonionic solute is dissolved in 36.30 g of benzene, C_6H_6 , the freezing point is 2.70 °C. Find the molar mass of this solute. The freezing point of benzene is 5.53 °C and the K_f is -5.12 °C /m.

10. A solution contains 21.6 g of a nonelectrolyte and 175 g of water. The water freezes at -7.18° C. Is the nonelectrolyte CH_3OH or C_2H_5OH? Explain your answer.

Colligative properties Worksheet

Part A-Calculation

1. Indicate how many particles are formed when the following solutes dissolve.

Solute	# of particles	Solute	# of particles
sucrose ($C_{12}H_{22}O_{11}$)		magnesium chloride ($MgCl_2$)	
sodium sulfate (Na_2SO_4)		Methanol (CH_3OH)	

2. When 5.0g of $CaCl_2$ dissolves in 50.0g of water, what is the boiling point of the solution? (K_b water= 0.512 °C/m)

3. Find the boiling point of a solution containing 6.0g benzene, C_6H_6, in 35g of naphthalene. (K_f water= 1.86 °C/m)

4. If you dissolve 26.0g of Epsom salt, $MgSO_4$, in 1.5kg of water, what is the freezing point of this solution? (K_f water= 1.86 °C/m)

Part B-Application

5. Salt is often used to remove ice from road. Explain how this process works in terms of colligative properties.

6. Which salt, NaCl or $CaCl_2$, has greater effect on freezing point? Explain.

7. How much would the boiling point of 1.0L of your pasta water change if you added 20g of NaCl to it?

8. Solution contains 21.6g of a nonelectrolytes and 175g of water. Water freezes at -7.18°C and K_f = 1.86°C/m. Is the nonelectrolytes CH_3OH or C_2H_5OH?

PRACTICE PROBLEMS ON NET IONIC EQUATIONS

Show the complete ionic and net ionic forms of the following equations. If all species are spectator ions, please indicate that no reaction takes place. Note: you need to make sure the original equation is balanced before proceeding! A set of solubility rules are given at the end of this document.

1. $AgNO_3(aq) + KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)$
2. $Mg(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow MgCO_3(s) + NaNO_3(aq)$
3. strontium bromide(aq) + potassium sulfate(aq) → strontium sulfate(s) + potassium bromide(aq)
4. manganese(II)chloride(aq) + ammonium carbonate(aq) → manganese(II)carbonate(s) + ammonium chloride(aq)
5. chromium(III)nitrate(aq) + iron(II)sulfate(aq) → chromium(III)sulfate(aq) + iron(II)nitrate(aq)

Please complete the following reactions, and show the complete ionic and net ionic forms of the equation:

6. $K_3PO_4(aq) + Al(NO_3)_3(aq) \rightarrow$
7. $BeI_2(aq) + Cu_2SO_4(aq) \rightarrow$
8. $Ni(NO_3)_3(aq) + KBr(aq) \rightarrow$
9. cobalt(III)bromide + potassium sulfide →
10. barium nitrate + ammonium phosphate →
11. calcium hydroxide + iron(III)chloride →
12. rubidium fluoride + copper(II)sulfate →

Solubility Rules

1. All salts of Group IA, and ammonium are soluble.
2. All salts of nitrates, chlorates and acetates are soluble.
3. All salts of halides are soluble except those of silver(I), copper(I), lead(II), and mercury(I).
4. All salts of sulfate are soluble except for barium sulfate, lead(II) sulfate, and strontium sulfate.
5. All salts of carbonate, phosphate and sulfite are insoluble, except for those of group IA and ammonium.
6. All oxides and hydroxides are insoluble except for those of group IA, calcium, strontium and barium.
7. All salts of sulfides and insoluble except for those of Group IA and IIA elements and of ammonium.

H-96 Name: ______________ period: ________

Net Ionic Equation Worksheet

Complete and balance each of the following equations carried out in aqueous solution. Also write the total ionic and net ionic equation for each.

(1) $BaBr_2(aq) + H_2SO_4(aq) \rightarrow$ $BaSO_{4(s)} + 2HBr$

$Ba^{+2}_{(aq)} + 2Br^{-}_{(aq)} + 2H^{+}_{(aq)} + SO_4^{-2}{}_{(aq)} \longrightarrow BaSO_{4(s)} + 2H^{+}_{(aq)} + 2Br^{-}_{(aq)}$

$Ba^{+2}_{(aq)} + SO_4^{-2}{}_{(aq)} \longrightarrow BaSO_{4(s)}$

(7) $KNO_3(aq) + LiC_2H_3O_2(aq) \rightarrow$

(13) $Na_2CO_3(aq) + BaCl_2(aq) \rightarrow$

(2) $Zn(s) + AgNO_3(aq) \rightarrow$

(8) $Zn(NO_3)_2(aq) + H_2S(aq) \rightarrow$

(14) $NaOH(aq) + HCl(aq) \rightarrow$

(3) $CaCl_2(aq) + Na_3PO_4(aq) \rightarrow$

(9) $AgNO_3(aq) + HCl(aq) \rightarrow$

Net Ionic Equation Worksheet

(15) $Cu(NO_3)_2(aq) + (NH_4)_2S(aq) \rightarrow$

(4) $Bi(OH)_3(s) + H_2SO_4(aq) \rightarrow$

(10) $Sr(NO_3)_2(aq) + H_2SO_4(aq) \rightarrow$

(16) $Pb(NO_3)_2(aq) + K_2CrO_4(aq) \rightarrow$

(5) $Pb(NO_3)_2(aq) + H_2S(aq) \rightarrow$

(11) $NaCl(aq) + KBr(aq) \rightarrow$

(17) $SrCl_2(aq) + AgNO_3(aq) \rightarrow$

(6) $Cu(OH)_2(s) + HBr(aq) \rightarrow$

(12) $KOH(aq) + HCl(aq) \rightarrow$ $HgSO_4(aq) + Mg(ClO_3)_2(aq) \rightarrow$

Name: ____________ Period: ______

Word Equations Worksheet

Write the word equations for each of the following chemical reactions:

1) When dissolved beryllium chloride reacts with dissolved silver nitrate in water, aqueous beryllium nitrate and silver chloride powder are made.

2) When isopropanol (C_3H_8O) burns in oxygen, carbon dioxide, water, and heat are produced.

3) When dissolved sodium hydroxide reacts with sulfuric acid (H_2SO_4), aqueous sodium sulfate, water, and heat are formed.

4) When fluorine gas is put into contact with calcium metal at high temperatures, calcium fluoride powder is created in an exothermic reaction.

5) When sodium metal reacts with iron (II) chloride, iron metal and sodium chloride are formed.

Honors Chemistry
Standard Heats of Formation

Name:

Heat of Formation Worksheet

Use a standard enthalpies of formation table to determine the change in enthalpy for each of these reactions.

a) $NaOH(s) + HCl(g) \longrightarrow NaCl(s) + H_2O(g)$

b) $2\ CO(g) + O_2(g) \longrightarrow 2\ CO_2(g)$

c) $CH_4(g) + 2\ O_2(g) \longrightarrow CO_2(g) + 2\ H_2O(l)$

d) $2\ H_2S(g) + 3\ O_2(g) \longrightarrow 2\ H_2O(l) + 2\ SO_2(g)$

e) $2\ NO(g) + O_2(g) \longrightarrow 2\ NO_2(g)$

Compound	ΔH_f (kJ/mol)	Compound	ΔH_f (kJ/mol)
$CH_4(g)$	-74.8	$HCl(g)$	-92.3
$CO_2(g)$	-393.5	$H_2O(g)$	-241.8
$NaCl(s)$	-411.0	$SO_2(g)$	-296.1
$H_2O(l)$	-285.8	$NH_4Cl(s)$	-315.4
$H_2S(g)$	-20.1	$NO(g)$	+90.4
$H_2SO_4(l)$	-811.3	$NO_2(g)$	+33.9
$MgSO_4(s)$	-1278.2	$SnCl_4(l)$	-545.2
$MnO(s)$	-384.9	$SnO(s)$	-286.2
$MnO_2(s)$	-519.7	$SnO_2(s)$	-580.7
$NaCl(s)$	-411.0	$SO_2(g)$	-296.1
$NaF(s)$	-569.0	$SO_3(g)$	-395.2
$NaOH(s)$	-426.7	$ZnO(s)$	-348.0
$NH_3(g)$	-46.2	$ZnS(s)$	-202.9

H-100

Hess's law: ΔH for a given equation is equal to the sum of the ΔH's for that equations steps

Chemistry 120
Hess's Law Worksheet

1. Calculate ΔH for the reaction C_2H_4 (g) + H_2 (g) → C_2H_6 (g), from the following data.

C_2H_4 (g) + 3 O_2 (g) → 2 CO_2 (g) + 2 H_2O (l)	ΔH = -1411. kJ/mole
C_2H_6 (g) + 7/2 O_2 (g) → 2 CO_2 (g) + 3 H_2O (l)	ΔH = -1560. kJ/mole
H_2 (g) + 1/2 O_2 (g) → H_2O (l)	ΔH = -285.8 kJ/mole

2. Calculate ΔH for the reaction 4 NH_3 (g) + 5 O_2 (g) → 4 NO (g) + 6 H_2O (g), from the following data.

N_2 (g) + O_2 (g) → 2NO (g)	ΔH = -180.5 kJ
N_2 (g) + 3 H_2 (g) → 2 NH_3 (g)	ΔH = -91.8 kJ
2 H_2 (g) + O_2 (g) → 2 H_2O (g)	ΔH = -483.6 kJ

3. Find ΔH_f^0 for acetic acid, $HC_2H_3O_2$, using the following thermochemical data.

$HC_2H_3O_2$ (l) + 2 O_2 (g) → 2 CO_2 (g) + 2 H_2O (l)	ΔH = -875. kJ/mole
C (s, graphite) + O_2 (g) → CO_2 (g)	ΔH = -394.51 kJ/mole
H_2 (g) + 1/2 O_2 (g) → H_2O (l)	ΔH = -285.8 kJ/mole

4. Calculate ΔH for the reaction CH_4 (g) + NH_3 (g) → HCN (g) + 3 H_2 (g), from the reactions.

N_2 (g) + 3 H_2 (g) → 2 NH_3 (g)	ΔH = -91.8 kJ
C (s, graphite) + 2 H_2 (g) → CH_4 (g)	ΔH = -74.9 kJ/mole
H_2 (g) + 2 C (s, graphite) + N_2 (g) → 2 HCN (g)	ΔH = +270.3 kJ

5. Calculate ΔH for the reaction 2 Al (s) + 3 Cl_2 (g) → 2 $AlCl_3$ (s) from the following data.

2 Al (s) + 6 HCl (aq) → 2 $AlCl_3$ (aq) + 3 H_2 (g)	ΔH = -1049. kJ
HCl (g) → HCl (aq)	ΔH = -74.8 kJ/mole
H_2 (g) + Cl_2 (g) → 2 HCl (g)	ΔH = -1845. kJ
$AlCl_3$ (s) → $AlCl_3$ (aq)	ΔH = -323. kJ/mole

Name: ______________________

Hess's Law Worksheet

1. Calculate ΔH for the reaction $C_2H_{4\,(g)} + H_{2(g)} \rightarrow C_2H_{6(g)}$, from the following data.

$C_2H_{4(g)} + 3\ O_{2(g)} \rightarrow 2\ CO_{2(g)} + 2\ H_2O_{(l)}$ ΔH= -1411.kJ/mole

$C_2H_{6(g)} + 7/2\ O_{2(g)} \rightarrow 2\ CO_{2(g)} + 3\ H_2O_{(l)}$ ΔH= -1560. kJ/mole

$H_{2(g)} + \frac{1}{2}\ O_{2(g)} \rightarrow H_2O_{(l)}$ ΔH= -285.8 kJ/mole

2. Calculate ΔH for the reaction $4\ NH_{3(g)} + 5\ O_{2(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$, from the following data.

$N_{2(g)} + O_{2(g)} \rightarrow 2\ NO_{(g)}$ ΔH= -180.5 k/J/mole

$N_{2(g)} + 3\ H_{2(g)} \rightarrow 2\ NH_{3(g)}$ ΔH= -91.8 kJ/mole

$2H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$ ΔH= -483.6 kJ/mole

3. Find ΔH_f^0 for acetic acid, $HC_2H_3O_2$, using the following thermochemical data.

$HC_2H_3O_{2(l)} + 2\ O_{2(g)} \rightarrow 2\ CO_{2(g)} + 2\ H_2O_{(l)}$ ΔH= -875. kJ/mole

$C_{(s,\ graphite)} + O_{2(g)} \rightarrow CO_{2(g)}$ ΔH= -394.51 kJ/mole

$H_{2(g)} + \frac{1}{2}\ O_{2(g)} \rightarrow H_2O_{(l)}$ ΔH= -285.8 kJ/mole

4. Calculate ΔH for the reaction $CH_{4(g)} + NH_{3(g)} \rightarrow HCN_{(g)} + 3\ H_{2(g)}$, from the following data.

$N_{2(g)} + 3\ H_{2(g)} \rightarrow 2\ NH_{3(g)}$ ΔH= -91.8 kJ/mole

$C_{(s,\ graphite)} + 2\ H_{2(g)} \rightarrow CH_{4(g)}$ ΔH= -74.9 kJ/mole

$H_{2(g)} + 2\ C_{(s,\ graphite)} + N_{2(g)} \rightarrow 2\ HCN_{(g)}$ ΔH= +270.3 kJ/mole

5. Calculate ΔH for the reaction $2Al_{(s)} + 3\ Cl_{2(g)} \rightarrow 2\ AlCl_{3(s)}$ from the following data.

$2\ Al_{(s)} + 6\ HCl_{(aq)} \rightarrow 2\ AlCl_{3(aq)} + 3\ H_{2(g)}$ ΔH= -1049. kJ/mole

$HCl_{(g)} \rightarrow HCl_{(aq)}$ ΔH= -74.8 kJ/mole

$H_{2(g)} + Cl_{2(g)} \rightarrow 2\ HCl_{(g)}$ ΔH= -1845. kJ/mole

$AlCl_{3(s)} \rightarrow AlCl_{3(aq)}$ ΔH= -323.kJ/mole

6H-2

Honors Chemistry
Standard Heats of Formation

Name:________________________________

Heat of Formation Worksheet

Use a standard enthalpies of formation table to determine the change in enthalpy for each of these reactions.

a) $NaOH(s) + HCl(g) \longrightarrow NaCl(s) + H_2O(g)$

b) $2\ CO(g) + O_2(g) \longrightarrow 2\ CO_2(g)$

c) $CH_4(g) + 2\ O_2(g) \longrightarrow CO_2(g) + 2\ H_2O(l)$

d) $2\ H_2S(g) + 3\ O_2(g) \longrightarrow 2\ H_2O(l) + 2\ SO_2(g)$

e) $2\ NO(g) + O_2(g) \longrightarrow 2\ NO_2(g)$

Compound	ΔH_f (kJ/mol)	Compound	ΔH_f (kJ/mol)
$CH_4(g)$	-74.8	$HCl(g)$	-92.3
$CO_2(g)$	-393.5	$H_2O(g)$	-241.8
$NaCl(s)$	-411.0	$SO_2(g)$	-296.1
$H_2O(l)$	-285.8	$NH_4Cl(s)$	-315.4
$H_2S(g)$	-20.1	$NO(g)$	+90.4
$H_2SO_4(l)$	-811.3	$NO_2(g)$	+33.9
$MgSO_4(s)$	-1278.2	$SnCl_4(l)$	-545.2
$MnO(s)$	-384.9	$SnO(s)$	-286.2
$MnO_2(s)$	-519.7	$SnO_2(s)$	-580.7
$NaCl(s)$	-411.0	$SO_2(g)$	-296.1
$NaF(s)$	-569.0	$SO_3(g)$	-395.2
$NaOH(s)$	-426.7	$ZnO(s)$	-348.0
$NH_3(g)$	-46.2	$ZnS(s)$	-202.9

$CO_{(g)}$ -99.0

6H-3

Bond Energy Calculations

Average Bond Enthalpies (kJ/mol)

Single Bonds

C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146		
C—N	293	N—O	201	O—F	190	Cl—F	253
C—O	358	N—F	272	O—Cl	203	Cl—Cl	242
C—F	485	N—Cl	200	O—I	234		
C—Cl	328	N—Br	243			Br—F	237
C—Br	276			S—H	339	Br—Cl	218
C—I	240	H—H	436	S—F	327	Br—Br	193
C—S	259	H—F	567	S—Cl	253		
		H—Cl	431	S—Br	218	I—Cl	208
Si—H	323	H—Br	366	S—S	266	I—Br	175
Si—Si	226	H—I	299			I—I	151
Si—C	301						
Si—O	368						

Multiple Bonds

C=C	614	N=N	418	O_2	495
C≡C	839	N≡N	941		
C=N	615			S=O	523
C≡N	891			S=S	418
C=O	799				
C≡O	1072				

1. Nitrogen + Hydrogen → Ammonia

2. Use bond energies to determine the energy change for the following reaction:
   $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

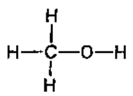
3. Use bond energies to determine the energy change for the following reaction:
   $C_2H_4(g) + F_2(g) \rightarrow C_2H_4F_2(g)$

4. Determine the energy change for the following reaction :
   methane + oxygen → carbon dioxide + water

6H-4

Bond Energy Calculations

5. Determine the enthalpy change for the following reaction:
   Hydrogen peroxide → water + oxygen

6. $C \equiv O$ + 2 H-H → 

7. H-H + 

6H-5

Name ____ GIBB'S FREE ENERGY PROBLEMS

1. The hydrogenation of ethene gas at 298. K shows a decrease in disorder ($\Delta S^\circ = -120.7$ J/(mol•K)) during an exothermic reaction ($\Delta H^\circ = -136.9$ kJ/mol). Determine whether the reaction is spontaneous or nonspontaneous by calculating ΔG°.
C_2H_4 (g) + H_2 (g) → C_2H_6 (g)

2. Copper (I) sulfide reacts with sulfur to produce copper (II) sulfide at 25°C. The process is exothermic ($\Delta H^\circ = -26.7$ kJ/mol) with a decrease in disorder ($\Delta S^\circ = -19.7$ J/(mol•K)). Determine the spontaneity of the reaction by calculating ΔG°.
Cu_2S (s) + S (s) → 2 CuS (s)

3. For a certain process at 300.0 K. $\Delta G = -77.0$ kJ/mol and $\Delta H = -56.9$ kJ/mol. Find the entropy change for this process.

4 The entropy of a system at 337 K increases by 221.7 J/mol•K. The free energy value is found to be –717.5 kJ/mol. Calculate the change in enthalpy of this system.

5 Determine if the following reaction is spontaneous or not 25°C by determining the free energy value.

$2\,NO_2(g) \rightarrow N_2O_4(g)$ ΔH= -57.2 kJ/mol ΔS = -175.9 J/ mol•K

6H-6

Writing Equilibrium Constant Expressions

1. Write the equilibrium constant expression for each of the following reactions.

   a. $2\ H_2O_2\ (g) \rightleftharpoons 2\ H_2O\ (g) + O_2\ (g)$

   b. $6\ H_2O_2\ (g) \rightleftharpoons 6\ H_2O\ (g) + 3\ O_2\ (g)$

   c. The reverse of the reaction in part a

   d. $2\ PbS\ (s) + 3\ O_2\ (g) \rightleftharpoons 2\ PbO\ (s) + 2\ SO_2\ (g)$

   e. $MgCl_2(s) \rightleftharpoons Mg^{2+}\ (aq) + 2\ Cl^-\ (aq)$

   f. The reverse of the reaction in part e

REDOX REACTIONS

Name ______________________

For the equations below, identify the substance oxidized, the substance reduced, the oxidizing agent, the reducing agent, and write the oxidation and reduction half reactions.

Example:

	oxidized		reduced		
	Mg	+	Br_2	$\rightarrow$	$MgBr_2$
	reducing agent		oxidizing agent		

oxidation half reaction: $Mg^{\circ} \rightarrow Mg^{+2} + 2e^-$

reduction half reaction: $2e^- + Br_2^{\circ} \rightarrow 2Br^-$

1. $2H_2 + O_2 \rightarrow 2H_2O$

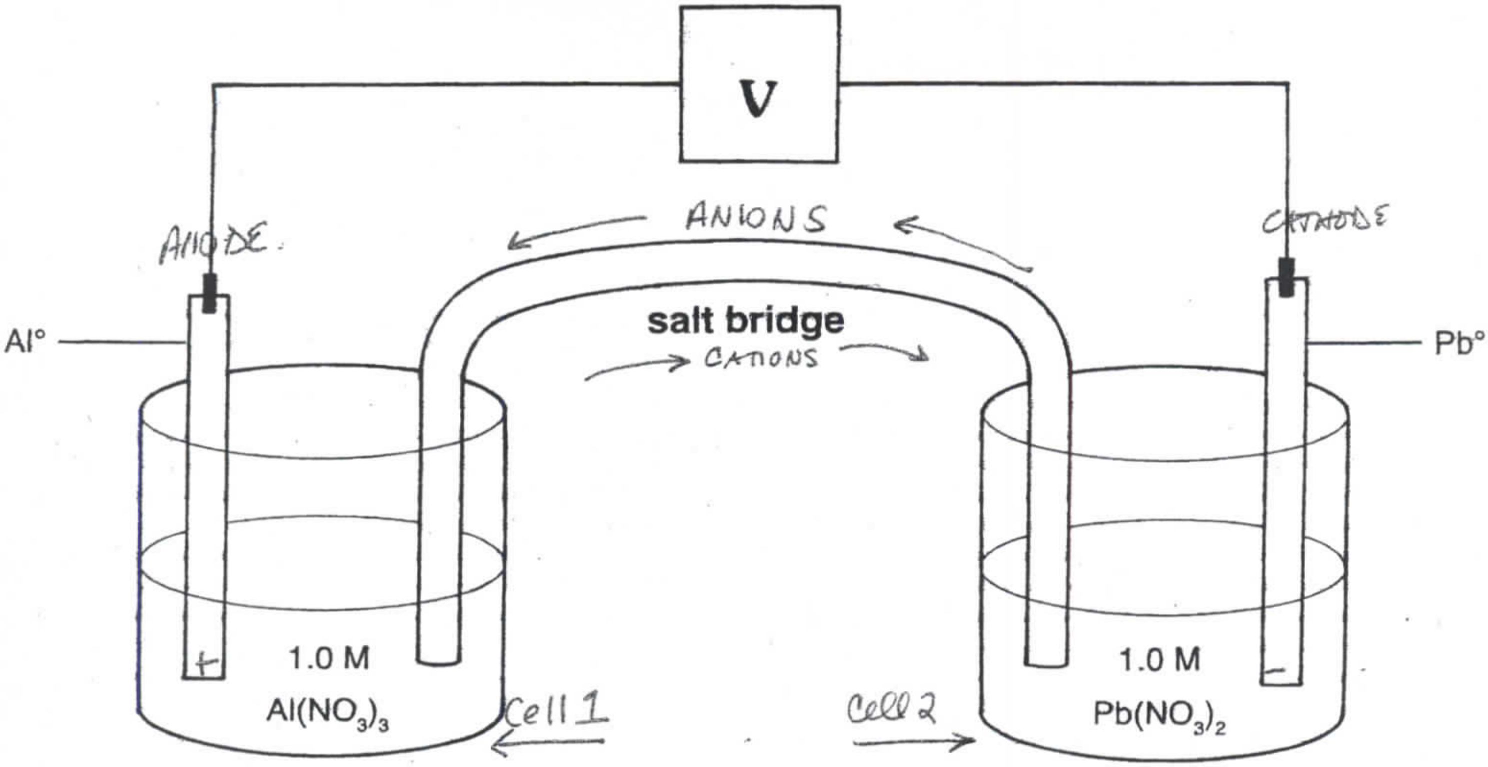
2. $Fe + Zn^{2+} \rightarrow Fe^{2+} + Zn$

3. $2Al + 3Fe^{+2} \rightarrow 2Al^{+3} + 3Fe$

4. $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$

THE ELECTROCHEMICAL CELL

Name ______________



Answer the questions below referring to the above diagram and a Table of Standard Electrode Potentials.

1. Which is more easily oxidized, metal, aluminum or lead? ______________
2. What is the balanced equation showing the spontaneous reaction that occurs?

   ____

3. ~~What is the maximum voltage that the above cell can produce?~~ disregard
4. What is the direction of electron flow in the wire? ______________
5. What is the direction of positive ion flow in the salt bridge? ______________
6. Which electrode is decreasing in size? ______________
7. Which electrode is increasing in size? ______________
8. What is happening to the concentration of aluminum ions? ______________
9. What is happening to the concentration of lead ions? ______________
10. What is the voltage in this cell when the reaction reaches equilibrium? ______________
11. Which is the anode? ______________
12. Which is the cathode? ______________
13. What is the positive electrode? ______________
14. What is the negative electrode? ______________

H111

Chapter 20 Worksheet: Redox

I. Determine what is oxidized and what is reduced in each reaction. Identify the oxidizing agent and the reducing agent, also.

1. $2Sr + O_2 \longrightarrow 2SrO$

2. $2Li + S \longrightarrow Li_2S$

3. $2Cs + Br_2 \longrightarrow 2CsBr$

4. $3Mg + N_2 \longrightarrow Mg_3N_2$

5. $4Fe + 3O_2 \longrightarrow 2Fe_2O_3$

6. $Cl_2 + 2NaBr \longrightarrow 2NaCl + Br_2$

7. $Si + 2F_2 \longrightarrow SiF_4$

8. $2Ca + O_2 \longrightarrow 2CaO$

9. $Mg + 2HCl \longrightarrow MgCl_2 + H_2$

10. $2Na + 2H_2O \longrightarrow 2NaOH + H_2$

11. Give the oxidation number of each kind of atom or ion.

a. sulfate b. Sn c. S^{2-} d. Fe^{3+} e. Sn^{4+} f. nitrate g. ammonium

12. Calculate the oxidation number of chromium in each of the following.

a. Cr_2O_3 b. $Na_2Cr_2O_7$ c. $CrSO_4$ d. chromate e. dichromate

13. Use the changes in oxidation numbers to determine which elements are oxidized and which are reduced in these reactions. (Note: it is not necessary to use balanced equations)

a. $C + H_2SO_4 \longrightarrow CO_2 + SO_2 + H_2O$

b. $HNO_3 + HI \longrightarrow NO + I_2 + H_2O$

c. $KMnO_4 + HCl \longrightarrow MnCl_2 + Cl_2 + H_2O + KCl$

d. $Sb + HNO_3 \longrightarrow Sb_2O_3 + NO + H_2O$

14. For each reaction in problem 13, identify the oxidizing agent and reducing agent.

15. Write half-reactions for the oxidation and reduction process for each of the following.

a. $Fe^{2+} + MnO_4^- \longrightarrow Fe^{3+} + Mn^{2+}$

b. $Sn^{2+} + IO_3^- \longrightarrow Sn^{4+} + I^-$

c. $S^{2-} + NO_3^- \longrightarrow S + NO$

d. $NH_3 + NO_2 \longrightarrow N_2 + H_2O$

16. Complet and balance each reaction using the half-reaction method.

a. $Fe^{2+} + MnO_4^- \longrightarrow Fe^{3+} + Mn^{2+}$

b. $Sn^{2+} + IO_3^- \longrightarrow Sn^{4+} + I^-$

c. $S^{2-} + NO_3^- \longrightarrow S + NO$

d. $NH_3 + NO_2 \longrightarrow N_2 + H_2O$

e. $Mn^{2+} + BiO_3^- \longrightarrow Bi^{2+} + MnO_4^-$

f. $I_2 + Na_2S_2O_3 \longrightarrow Na_2S_2O_4 + NaI$

Worksheet 7 - Oxidation/Reduction Reactions

Oxidation number rules:

Elements have an oxidation number of **0**
Group I and **II** – In addition to the elemental oxidation state of 0, Group I has an oxidation state of **+1** and Group II has an oxidation state of **+2**.
Hydrogen –usually **+1**, except when bonded to Group I or Group II, when it forms hydrides, **-1**.
Oxygen – usually **-2**, except when it forms a O-O single bond, a peroxide, when it is **-1**.
Fluorine is always **-1**. Other halogens are usually -1, except when bonded to O.

1. Assign **oxidation numbers** to each of the atoms in the following compounds:

Na_2CrO_4	Na =	O =	Cr =
$K_2Cr_2O_7$	K =	O =	Cr =
CO_2	O =	C =	
CH_4	H =	C =	
$HClO_4$	O =	H =	Cl =
MnO_2	O =	Mn =	
SO_3^{2-}	O =	S =	
SF_4	F =	S =	

a. What is the range of oxidation states for **carbon**?

b. Which compound has C in a +4 state?

c. Which compound has C in a -4 state?

2. Nitrogen has 5 valence electrons (Group V). It can gain up to 3 electrons (-3), or lose up to 5 (+5) electrons. Fill in the missing names or formulas and assign an oxidation state to each of the following nitrogen containing compounds:

name	**formula**	**oxidation state of N**
	NH_3	
nitrogen		
nitrite		
	NO_3^-	
dinitrogen monoxide		
	NO_2	
hydroxylamine	NH_2OH	
nitrogen monoxide		
hydrazine	N_2H_4	

During chemical reactions, the **oxidation state** of atoms can change. This occurs when compounds gain or lose electrons, or when the **bonds** to an atom change. This is illustrated by the reaction between nitrogen and hydrogen to make ammonia:

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

a. Assign **oxidation numbers** to each of the atoms in this reaction.

N (in N_2) = N (in NH_3) =

H (in H_2) = H (in NH_3) =

When an oxidation number **increases,** that species has been **oxidized**.

b. Which reactant undergoes an increase in its oxidation number?

When an oxidation number **decreases,** that species has been **reduced**.

c. Which reactant undergoes a decrease in its oxidation number?

The species that is oxidized is called the **reducing agent** because it gives up an electron, so that another species can gain an electron (be reduced).

d. What is the **reducing agent** in this reaction?

The species that is reduced is called the **oxidizing agent** because it takes an electron away from another group, raising that group's oxidation number.

e. What is the **oxidizing agent** in this reaction?

3. In each of the following reactions, assign **oxidation numbers** to all of the elements and identify the **oxidizing** and **reducing agents** and the **change in oxidation number**.

a. $4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$

change in oxidation number

oxidizing agent

reducing agent

b. $Cr_2O_7^{2-} + 2OH^- \rightarrow 2\ CrO_4^{2-} + H_2O$

change in oxidation number

oxidizing agent

reducing agent

c. $NH_4NO_2 \rightarrow N_2 + 2\ H_2O$

change in oxidation number

oxidizing agent

reducing agent

d. $P_4 + 10\ Cl_2 \rightarrow 4\ PCl_5$

change in oxidation number

oxidizing agent

reducing agent

e. $2\ Cr^{3+} + H_2O + 6\ ClO_3^- \rightarrow Cr_2O_7^{2-} + 6ClO_2 + 2\ H^+$

change in oxidation number

oxidizing agent

reducing agent

Balancing Redox Reactions

Oxidation/Reduction (Redox) reactions can be balanced using the oxidation state changes, as seen in the previous example. However, there is an easier method, which involves breaking a redox reaction into two **half- reactions**. This is best shown by working an example.

Hydrobromic acid will react with permanganate to form elemental bromine and the manganese(II) ion. The unbalanced, net reaction is shown below,

$$\mathbf{Br^- + MnO_4^- \rightarrow Br_2 + Mn^{2+}}$$

1. Break this into two **half-reactions**, one involving bromine and the other involving manganese.

Bromine half-reaction	**Manganese half-reaction**
$Br^- \rightarrow Br_2$	$MnO_4^- \rightarrow Mn^{2+}$

2. First balance the bromine half-reaction first.

   a. Balance the **bromine** atoms of the reaction

   $$\_\_\_ Br^- \rightarrow \_\_\_ Br_2$$

   b. Now balance **charge** by adding **electrons** (**e**$^-$)

   $$\_\_\_ Br^- \rightarrow \_\_\_ Br_2$$

   This half-reaction is **producing/consuming electrons**. This is an **oxidation/reduction** half-reaction. Confirm this by assigning oxidation numbers to the bromine species.

3. Next, balance the manganese half-reaction.

   a. Balance the **manganese** atoms of the half-reaction

   $$\_\_\_ MnO_4^- \rightarrow \_\_\_ Mn^{2+}$$

   b. Next, balance **oxygen** by adding water molecules (**H_2O**)

   $$\_\_\_ MnO_4^- \rightarrow \_\_\_ Mn^{2+}$$

The overall balanced reaction under basic conditions is:

$$\_\_\_Br^- + \_\_\_MnO_4^- + \_\_\_H_2O \rightarrow \_\_\_Br_2 + \_\_\_Mn^{2+} + \_\_\_OH^-$$

Now, balance the redox reaction between methanol and dichromate, which produces methanal and chromium (III), as shown below:

$$\mathbf{CH_3OH + Cr_2O_7^{2-} \rightarrow CH_2O + Cr^{3+}}$$

First, separate this into two half-reactions
Then, balance the redox active species.
Then, balance oxygens with H_2O
Balance hydrogen with H^+
Balance charge with electrons.
Equalize the number of electrons lost and gained

This indicates that the reaction must be carried out in an **acidic** solution.

To carry it out in a **basic** solution, just add enough $\mathbf{OH^-}$ to neutralize the acid, $\mathbf{H^+}$

CHEM1101 Worksheet 12: Electrochemistry

Model 1: Reduction Potentials

The **standard reduction potential**, E^0_{red} has units of volts (V) and is a measure of a species ability to attract electrons. The *more positive* the reduction potential, the *stronger* is the attraction for electrons. Put another way, the *more positive* the reduction potential, the easier the reduction occurs. Some standard reduction potentials are given below.

	Reduction reaction	E^0_{red} (V)
(1)	$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
(2)	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.34
(3)	$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$	-0.14
(4)	$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	-0.76

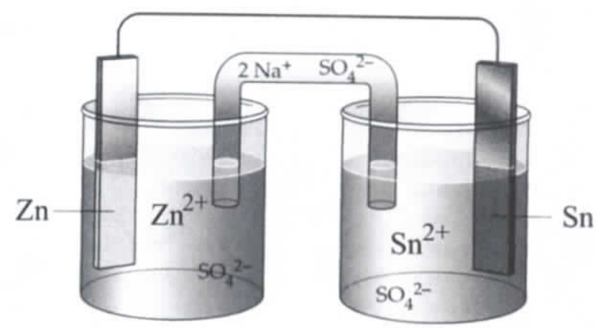
Critical thinking questions

1. Which is a stronger **oxidising** agent: Ag^+ or Cu^{2+}? Explain how you can tell in terms of the reduction potentials.

2. If reactions (1) and (2) are added together as a redox reaction which do you think will proceed as a reduction and which as an oxidation? (*Hint*: which one will reverse?)

3. Apply the same logic to reactions (3) and (4). Does it matter that they both have negative reduction potentials?

Model 2: Voltaic Cells



We can harness the electrical energy in a redox reaction, to make a battery, by setting up a **voltaic cell**. To do this, two half reactions are separated into compartments and electrodes are used to facilitate the electron transfer. The potentials for the two reactions are:

$$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s) \qquad E^0_{red} = -0.14 \text{ V}$$
$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^- \qquad E^0_{ox} = +0.76 \text{ V}$$

This gives an overall voltmeter reading of:

$$E^0_{cell} = E^0_{ox} + E^0_{red} = +0.62 \text{ V}.$$

Critical thinking questions

1. Explain why the Zn half reaction is proceeding as an oxidation and why +0.76 V is used as the potential for its half cell instead of -0.76 V as in the table in Model 1?

2. Which electrode (Zn or Sn) will *lose* mass and which one will *gain* mass?

3. Does oxidation or reduction occur at the cathode?

4. Which of the following statements are correct?

   (a) Electrons flow through the wire, towards the zinc electrode.
   (b) Electrons flow through the wire, towards the tin electrode.
   (c) Electrons flow through the salt bridge, towards the zinc electrode.
   (d) Electrons flow through the salt bridge, towards the tin electrode.

5. Electrons flow from the *negative* electrode to the *positive* electrode. Which is positive, the anode or the cathode?

6. The salt bridge contains $Na^+(aq)$ and $SO_4^{2-}(aq)$. Do these ions *move* when the cell is operating and, if so, in which direction(s)?

7. If an electrochemical cell with Ag and Cu electrodes was setup, what would be the two half reactions, which would be the cathode and which would be the anode, and what would be the standard cell potential? (*Hint*: use the standard reduction potentials in Model 1.)

8. If an electrochemical cell with Sn and Cu electrodes was setup, what would be the two half reactions, which would be the cathode and which would be the anode, and what would be the standard cell potential?

9. If an electrochemical cell with Sn and Zn electrodes was setup, what would be thetwo half reactions, which would be the cathode and which would be the anode, and what would be the standard cell potential?

10. Which combination of the half cells in Table 1 would make the highest voltage battery?

11. Nicotine adenine dinucleotide (NAD) is involved in redox chemistry throughout the respiratory system. The reduced form of NAD is written as NADH and the oxidised form is written as NAD^+. The standard reduction reaction and potential of NAD is given by:

$$NAD^+ + H^+ + 2e^- \rightarrow NADH \quad E° = -0.105\ V$$

NAD is combined with each of the following reactions:

(a) $CO_2 + H^+ + 2e^- \rightarrow HCOO^-$ $E° = -0.20\ V$
(b) $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ $E° = +0.82\ V$

Write the overall reaction for each of the cells in the direction of spontaneous change. Is the NAD reduced or oxidised in these reactions?

Galvanic (Voltaic) Cells Problems

Name______________________________ Hr_____

1. Regarding the following reaction:

$$F_{2\,(g)} + 2\,I^-_{(aq)} \rightarrow 2\,F^-_{(aq)} + I_{2\,(s)}$$

a. List the species being oxidized:______ List the species being reduced:_____

b. Calculate E° for this cell.

c. Which species receives electrons from the anode?_____

d. Which species donates electrons to the cathode?______

2. Given the following half-reactions:

$Co^{+2} + 2e^- \rightarrow Co$ $E^\circ = -0.277\ V$

$Ce^{+4} + e^- \rightarrow Ce^{+3}$ $E^\circ = 1.61\ V$

a. Write the overall equation for the galvanic cell. Be sure to balance the number of electrons.

b. Calculate the E° for the cell.

c. Designate which ½ reaction occurs...

at the anode: ______________________ and at the cathode: ______________________

d. Describe the direction of electron flow.

3. Find the following using a voltaic cell with the line notation: $Al_{(s)}|Al^{3+}||Pb^{2+}|Pb_{(s)}$

a. Calculate the E° for the reaction.

b. Write the balanced overall REDOX equation for the galvanic cell.

4. Using standard reduction potentials, calculate the cell potential (E°) for each of the following reactions:

a. $H_{2\,(g)} + I_{2\,(s)} \rightarrow 2H^+_{(aq)} + 2I^-_{(aq)}$

b. $Ni_{(s)} + 2Ce^{+4}_{(aq)} \rightarrow Ni^{+2}_{(aq)} + 2Ce^{+3}_{(aq)}$

5. For the following voltaic cell:

$$Fe_{(s)} \mid 1.00\ M\ Fe^{2+}_{(aq)} \parallel 1.00\ M\ Cu^{2+}_{(aq)} \mid Cu_{(s)}$$

(a). Draw a picture of the cell. Include a salt bridge in your picture.
(b). Write the cell half-reactions.
(b). Label one half-reaction as the oxidation, the other as the reduction.
(c). Label on half-reaction as the cathode, the other the anode.
(d). Label the direction of electron flow, and label one electrode as the cathode and the other electrode as the anode.
(e). Write the overall cell reaction.
(f). Determine the standard voltage (E^o) for the battery

6. Find the following using a voltaic cell containing $Cr_{(s)}/Cr^{3+}_{(aq)}$ along with $Co_{(s)}/Co^{2+}_{(aq)}$

a. Write the overall equation for the galvanic cell and calculate the E^o for the cell.

b. Draw and describe all parts of the cell.

Balancing Redox Equations by the Ion-Electron Method

There are two principal methods for balancing redox equations:

- oxidation state method
- ion-electron method.

The latter is easier to use with redox reactions in aqueous solution and if necessary can be adapted to many situations that are not in aqueous solution. Our primary interest will be in aqueous-solution redox; therefore, we will use the ion-electron method. One of the major advantages of this method is that it makes it completely unnecessary to assign individual oxidation numbers.

To balance a redox equation by the ion-electron method, carry out the following steps in this sequence:

1. **Separate the skeletal equation into two half reactions.** One half reaction will be a reduction and the other will be an oxidation. It is not necessary at this stage to identify which is which.
2. **Balance each half reaction separately.** Balance atoms on each side of a half reaction by inspection. If the reaction occurs in acidic medium, you may add H_2O and/or H^+ to balance oxygen and/or hydrogen. If the reaction occurs in basic medium, you may add H_2O and/or OH^- to balance oxygen and/or hydrogen. Do *not* add any other new species (e.g., O_2, H_2) unless already a part of the skeletal half reaction.
3. **Balance the net charge across each half reaction by adding electrons to the side with the more positive net ionic charge.** If by this process electrons are added on the left side of a half reaction, the half reaction is a reduction. If electrons are added to the right side, the half reaction is an oxidation. (If you add electrons to the same side in both half reactions, something is wrong!)
4. **Multiply both half-reactions by appropriate whole number factors, so that the number of electrons is the same in both half reactions and will cancel when the two are added together.**
5. **Add the two multiplied half reactions together to obtain the overall redox equation.**
6. **Check the balance.** No electrons should appear in the overall redox equation. Not only should there be an element-by-element balance across the equation, but also the net charge (the sum of both ionic charges and electron charges) on both sides of the equation should be equal.

Note that this procedure does not involve assigning oxidation numbers. Nonetheless, if oxidation numbers are assigned to the balanced equation, it will always occur that the reduction involves lowering an oxidation state of some element, and the oxidation involves raising an oxidation state of some element. The following examples illustrate the ion-electron procedure, starting from the skeletal equation in either acidic or basic solution.

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Balancing Redox Reactions 2:
The Ion-Electron Method

In the first redox reaction worksheet, we saw the oxidation number method of balancing equations. This worksheet shows you another method.

The steps for balancing a redox reaction using the *ion-electron method* are:

[1] **Break the equation into two half-reactions**, one for the oxidation step (loss of electrons) and one for the reduction step (gain of electrons). You will still need to use oxidation numbers to know which is which.

[2] Obtain material balance (i.e. **balance the atoms**) in each half-reaction.
- [a] Balance everything other than hydrogen and oxygen.
- [b] Balance oxygen by adding H_2O to the other side.
- [c] Balance hydrogen by adding H^+ to the other side.
- [d] IF THE REACTION IS IN BASIC SOLUTION, add equal amounts of OH^- to both sides to neutralize the H^+. The OH^- and H^+ combine to form water and leave excess OH^- on the other side. Cancel any water that appears on both sides. (ignore step d if solution is acidic)

[3] Obtain charge balance for each half-reaction by **adding electrons as a product/reactant to the more positive side**.

[4] **Combine the half-reactions to cancel the electrons.** You may have to multiply the equations by whole numbers to do this.

Example 1: Balance the following redox reaction using the ion-electron method:

$$NO + SO_4^{2-} \rightarrow NO_3^- + SO_2$$

Solution: Following the steps above:

[1] Nitrogen gets oxidized, and sulphur is reduced, so the half-reactions are:

$$NO \rightarrow NO_3^- \qquad SO_4^{2-} \rightarrow SO_2$$

[2] We balance the atoms:

$$NO + 2\,H_2O \rightarrow NO_3^- \qquad SO_4^{2-} \rightarrow SO_2 + 2\,H_2O$$
$$NO + 2\,H_2O \rightarrow NO_3^- + 4\,H^+ \qquad 4\,H^+ + SO_4^{2-} \rightarrow SO_2 + 2\,H_2O$$

[3] We add electrons so that the charge balances:

$$NO + 2\,H_2O \rightarrow NO_3^- + 4\,H^+ + 3\,e^- \qquad 4\,H^+ + SO_4^{2-} + 2\,e^- \rightarrow SO_2 + 2\,H_2O$$

[4] And finally we cancel the electrons:

$$2\,NO \cancel{+ 4\,H_2O} \rightarrow 2\,NO_3^- \cancel{+ 8\,H^+ + 6\,e^-}$$
$$^{4}\cancel{12}\,H^+ + 3\,SO_4^{2-} \cancel{+ 6\,e^-} \rightarrow 3\,SO_2 + \cancel{6}^{2}\,H_2O$$
$$\overline{2\,NO + 4\,H^+ + 3\,SO_4^{2-} \rightarrow 2\,NO_3^- + 3\,SO_2 + 2\,H_2O}$$

Authored by Gordon Wong

[5] With sulphates and nitrates, this reaction is not likely to take place in basic solution, but if it were, we would not be able to have H^+ in the final equation. We would add OH^- to both sides to cancel the H^+ that is there:

$$\begin{aligned}
2\,NO + 4\,H^+ + 3\,SO_4^{2-} &\rightarrow 2\,NO_3^- + 3\,SO_2 + 2\,H_2O \\
2\,NO + 4\,H^+ + 4\,OH^- + 3\,SO_4^{2-} &\rightarrow 2\,NO_3^- + 3\,SO_2 + 2\,H_2O + 4\,OH^- \\
2\,NO + {}^{2}\cancel{4}\,H_2O + 3\,SO_4^{2-} &\rightarrow 2\,NO_3^- + 3\,SO_2 \cancel{+ 2\,H_2O} + 4\,OH^- \\
2\,NO + 2\,H_2O + 3\,SO_4^{2-} &\rightarrow 2\,NO_3^- + 3\,SO_2 + 4\,OH^-
\end{aligned}$$

EXERCISES

A. For each redox equation, determine (a) the oxidation half-reaction, (b) the reduction half-reaction, and (c) the balanced redox reaction.

1) $Fe^{3+} + Sn^{2+} \rightarrow Fe^{2+} + Sn^{4+}$

2) $Fe^{2+} + C\ell O_3^- + H^+ \rightarrow Fe^{3+} + C\ell^- + H_2O$

3) $Cu + NO_3^- + H^+ \rightarrow Cu^{2+} + NO + H_2O$

4) $S_2O_4^{2-} + Ag_2O + H_2O \rightarrow SO_3^{2-} + Ag + H^+$

5) $MnO_4^- + C\ell^- + H^+ \rightarrow Mn^{2+} + C\ell_2 + H_2O$

6) $MnO_4^- + S^{2-} + H_2O \rightarrow MnO_2 + S + OH^-$

7) $Cr_2O_7^{2-} + I^- \rightarrow Cr^{3+} + I_2$ *acidic solution

8) $MnO_4^- + Sn \rightarrow MnO_2 + Sn(OH)_3^-$ *basic solution

9) $H_2O_2 \rightarrow H_2O + O_2$

SOLUTIONS

A. (1)a) $Sn^{2+} \rightarrow Sn^{4+} + 2\ e^-$ (b) $Fe^{3+} + e^- \rightarrow Fe^{2+}$
(c) a + 2b: $2\ Fe^{3+} + Sn^{2+} \rightarrow 2\ Fe^{2+} + Sn^{4+}$
(2)a) $Fe^{2+} \rightarrow Fe^{3+} + e^-$ (b) $C\ell O_3^- + 6\ H^+ + 6\ e^- \rightarrow C\ell^- + 3\ H_2O$
(c) 6a + b: $6\ Fe^{2+} + C\ell O_3^- + 6\ H^+ \rightarrow 6\ Fe^{3+} + C\ell^- + 3\ H_2O$
(3)a) $Cu \rightarrow Cu^{2+} + 2\ e^-$ (b) $NO_3^- + 4\ H^+ + 3\ e^- \rightarrow NO + 2\ H_2O$
(c) 3a + 2b: $3\ Cu + 2\ NO_3^- + 8\ H^+ \rightarrow 3\ Cu^{2+} + 2\ NO + 4\ H_2O$
(4)a) $S_2O_4^{2-} + 2\ H_2O \rightarrow 2\ SO_3^{2-} + 4\ H^+ + 2\ e^-$ (b) $Ag_2O + 2H^+ + 2\ e^- \rightarrow 2\ Ag + H_2O$
(c) a + b: $S_2O_4^{2-} + Ag_2O + H_2O \rightarrow 2\ SO_3^{2-} + 2\ Ag + 2\ H^+$
(5)a) $2\ C\ell^- \rightarrow C\ell_2 + 2\ e^-$ (b) $MnO_4^- + 5\ e^- + 8\ H^+ \rightarrow Mn^{2+} + 4\ H_2O$
(c) 5a + 2b: $2\ MnO_4^- + 10\ C\ell^- + 16\ H^+ \rightarrow 2\ Mn^{2+} + 5\ C\ell_2 + 8\ H_2O$
(6)a) $S^{2-} \rightarrow S + 2\ e^-$ (b) $MnO_4^- + 2\ H_2O + 3\ e^- \rightarrow MnO_2 + 4\ OH^-$
(c) 3a + 2b: $2\ MnO_4^- + 3\ S^{2-} + 4\ H_2O \rightarrow 2\ MnO_2 + 3\ S + 8\ OH^-$
(7)a) $2\ I^- \rightarrow I_2 + 2\ e^-$ (b) $Cr_2O_7^{2-} + 6\ e^- + 14\ H^+ \rightarrow 2\ Cr^{3+} + 7\ H_2O$
(c) 3a + b: $Cr_2O_7^{2-} + 6\ I^- + 14\ H^+ \rightarrow 2\ Cr^{3+} + 3\ I_2 + 7\ H_2O$
(8)a) $Sn + 3\ OH^- \rightarrow Sn(OH)_3^- + 2\ e^-$ (b) $MnO_4^- + 2\ H_2O + 3\ e^- \rightarrow MnO_2 + 4\ OH^-$
(c) 3a + 2b: $2\ MnO_4^- + 3\ Sn + OH^- + 4\ H_2O \rightarrow 2\ MnO_2 + 3\ Sn(OH)_3^-$
(9)a) $H_2O_2 \rightarrow O_2 + 2\ H^+ + 2\ e^-$ (b) $H_2O_2 + 2\ H^+ + 2\ e^- \rightarrow 2\ H_2O$
(c) a + b: $2\ H_2O_2 \rightarrow 2\ H_2O + O_2$

Balancing Redox Reactions Worksheet 1

Balance each redox reaction in acid solution.

$Mn^{2+} + BiO_3^- \rightarrow MnO_4^- + Bi^{3+}$

$MnO_4^- + S_2O_3^{2-} \rightarrow S_4O_6^{2-} + Mn^{2+}$

$ClO_3^- + Cl^- \rightarrow Cl_2 + ClO_2$

$P + Cu^{2+} \rightarrow Cu + H_2PO_4^-$

$PH_3 + I_2 \rightarrow H_3PO_2^- + I^-$

$NO_2 \rightarrow NO_3^- + NO$

Basic Solutions

$MnO_4^- + C_2O_4^{2-} \rightarrow MnO_2 + CO_2$

$ClO_2 \rightarrow ClO_2^- + ClO_3^-$

$Cu(NH_3)_4^{2+} + S_2O_4^{2-} \rightarrow SO_3^{2-} + Cu + NH_3$

$Zn + NO_3^- \rightarrow Zn(OH)_4^{2-} + NH_3$

$Al + OH^- \rightarrow AlO_2^- + H_2$

$Zn \rightarrow Zn(OH)_4^{2-} + H_2$

Answers
Acidic:

$14\ H^+ + 2\ Mn^{2+} + 5\ BiO_3^- \rightarrow 2\ MnO_4^- + 5\ Bi^{3+} + 7\ H_2O$

$16\ H^+ + 2\ MnO_4^- + 10\ S_2O_3^{2-} \rightarrow 2\ Mn^{2+} + 8\ H_2O + 5\ S_4O_6^{2-}$

$4\ H^+ + 2\ ClO_3^- + 2\ Cl^- \rightarrow 2\ ClO_2 + 2\ H_2O + Cl_2$

$8\ H_2O + 2\ P + 5\ Cu^{2+} \rightarrow 2\ H_2PO_4^- + 12\ H^+ + 5\ Cu)$

$4\ H_2O + 2\ PH_3 + 3\ I_2 \rightarrow 2\ H_3PO_2^- + 8\ H^+ + 6\ I^-$

$H_2O + 3\ NO_2 \rightarrow 2\ NO_3^- + NO + 2\ H^+$

Basic:

$4\ H_2O + 2\ MnO_4^- + 3\ C_2O_4^{2-} \rightarrow 2\ MnO_2 + 8\ OH^- + 6\ CO_2$

$2\ OH^- + 2\ ClO_2 \rightarrow ClO_2^- + ClO_3^- + H_2O$

$Cu(NH_3)_4^{2+} + S_2O_4^{2-} + 4\ OH^- \rightarrow Cu + 4\ NH_3 + 2\ SO_3^{2-} + 2\ H_2O$

$6\ H_2O + NO_3^- + 7\ OH^- + 4\ Zn \rightarrow NH_3 + 4\ Zn(OH)_4^{2-}$

$2\ Al + 2\ H_2O + 2\ OH^- \rightarrow 2\ AlO_2^- + 3\ H_2$

$2\ OH^- + Zn + 2\ H_2O \rightarrow Zn(OH)_4^{2-} + H_2$

Balancing Redox Reactions Worksheet 2

Balance each redox reaction in acid solution.

$H_2O_2 + Cr_2O_7^{2-} \rightarrow O_2 + Cr^{3+}$

$TeO_3^{2-} + N_2O_4 \rightarrow Te + NO_3^-$

$ReO_4^- + IO^- \rightarrow IO_3^- + Re$

$PbO_2 + I_2 \rightarrow Pb^{2+} + IO_3^-$

$As \rightarrow H_2AsO_4^- + AsH_3$

Balance each redox reaction in basic solution.

$O_2 + Cr^{3+} \rightarrow H_2O + Cr_2O_7^{2-}$

$Te + NO_3^- \rightarrow TeO_3^{2-} + N_2O_4$

$IO_3^- + Re \rightarrow ReO_4^- + IO^-$

$Pb^{2+} + IO_3^- \rightarrow PbO_2 + I_2$

$Cr_2O_7^{2-} + Hg \rightarrow Hg^{2+} + Cr^{3+}$

Answers:

$8H^+ + 3H_2O_2 + Cr_2O_7^{2-} \rightarrow 3O_2 + 2Cr^{3+} + 7H_2O$

$TeO_3^{2-} + 2N_2O_4 + H_2O \rightarrow Te + 4NO_3^- + 2H^+$

$4H^+ + 4ReO_4^- + 7IO^- \rightarrow 7IO_3^- + 4Re + 2H_2O$

$8H^+ + 5PbO_2 + I_2 \rightarrow 5Pb^{2+} + 2IO_3^- + 4H_2O$

$12H_2O + 8As \rightarrow 3H_2AsO_4^- + 5AsH_3 + 3H^+$

$3O_2 + 8OH^- + 2Cr^{3+} \rightarrow H_2O + 3H_2O_2 + Cr_2O_7^{2-}$

$H_2O + Te + 4NO_3^- \rightarrow TeO_3^{2-} + 2OH^- + 2N_2O_4$

$7IO_3^- + 4OH^- + 4Re \rightarrow 4ReO_4^- + 7IO^- + 2H_2O$

$8OH^- + 5Pb^{2+} + 2IO_3^- \rightarrow 5PbO_2 + I_2 + 4H_2O$

$7H_2O + Cr_2O_7^{2-} + 3Hg \rightarrow 3Hg^{2+} + 14OH^- + 2Cr^{3+}$

Standard Reduction Potentials

Reduction Half-Reaction	Standard Reduction Potential (V)
$F_2(g)+2e^- \rightarrow 2F^-(aq)$	+2.87
$S_2O_8^{2-}(aq)+2e^- \rightarrow 2SO_4^{2-}(aq)$	+2.01
$O_2(g)+4H^+(aq)+4e^- \rightarrow 2H_2O(l)$	+1.23
$Br_2(l)+2e^- \rightarrow 2Br^-(aq)$	+1.09
$Ag^+(aq)+e^- \rightarrow Ag(s)$	+0.80
$Fe^{3+}(aq)+e^- \rightarrow Fe^{2+}(aq)$	+0.77
$I_2(l)+2e^- \rightarrow 2I^+(aq)$	+0.54
$Cu^{2+}(aq)+2e^- \rightarrow Cu(s)$	+0.34
$Sn^{4+}(aq)+2e^- \rightarrow Sn^{2+}(aq)$	+0.15
$S(s)+2H^+(aq)+2e^- \rightarrow H_2S(g)$	+0.14
$2H^+(aq)+2e^- \rightarrow H_2(g)$	**0.00**
$Sn^{2+}(aq)+2e^- \rightarrow Sn(g)$	-0.14
$V^{3+}(aq)+e^- \rightarrow V^{2+}(aq)$	-0.26
$Fe^{2+}(aq)+2e^- \rightarrow Fe(s)$	-0.44
$Cr^{3+}(aq)+3e^- \rightarrow Cr(s)$	-0.74
$Zn^{2+}(aq)+2e^- \rightarrow Zn(s)$	-0.76
$Mn^{2+}(aq)+2e^- \rightarrow Mn(s)$	-1.18
$Na^+(aq)+e^- \rightarrow Na(s)$	-2.71
$Li^+(aq)+e^- \rightarrow Li(s)$	-3.04

Calculating pH and pOH worksheet

1) What is the pH of a 0.0235 M HCl solution?

2) What is the pOH of a 0.0235 M HCl solution?

3) What is the pH of a 6.50×10^{-3} M KOH solution? (Hint: this is a basic solution – concentration is of OH^-)

4) A solution is created by measuring 3.60×10^{-3} moles of NaOH and 5.95×10^{-4} moles of HCl into a container and then water is added until the final volume is 1.00 L. What is the pH of this solution?

5) What is the pH of a 6.2×10^{-5} M NaOH solution? (Hint: this is a basic solution – concentration is of OH^-)

6) A solution with a H^+ concentration of 1.00×10^{-7} M is said to be neutral. Why?

Strong Acids and Bases

7 STRONG ACIDS

HCl - hydrochloric acid

HBr - hydrobromic acid

HI - hydroiodic acid

HNO_3 - nitric acid

$HClO_3$ - chloric acid

$HClO_4$ - perchloric acid

H_2SO_4 - sulfuric acid

8 STRONG BASES

$LiOH$ - lithium hydroxide

$NaOH$ - sodium hydroxide

KOH - potassium hydroxide

$RbOH$ - rubidium hydroxide

$CsOH$ - cesium hydroxide

$Ca(OH)_2$ - calcium hydroxide

$Sr(OH)_2$ - strontium hydroxide

$Ba(OH)_2$ - barium hydroxide

Buffers

What Is a Buffer?

A buffer is an aqueous solution that has a highly stable pH. If you add an acid or a base to a buffered solution, its pH will not change significantly. Similarly, adding water to a buffer or allowing water to evaporate will not change the pH of a buffer.

How Do You Make a Buffer?

A buffer is made by mixing a large volume of a weak acid or weak base together with its conjugate. A weak acid and its conjugate base can remain in solution without neutralizing each other. The same is true for a weak base and its conjugate acid.

How Do Buffers Work?

When hydrogen ions are added to a buffer, they will be neutralized by the base in the buffer. Hydroxide ions will be neutralized by the acid. These neutralization reactions will not have much effect on the overall pH of the buffer solution.