General Chemistry I
Answers to Practice Problems, Chapters 4, 6, and 7

1.   a) sodium sulfate   b) nitrous acid (<u>not</u> "hydrogen nitrite")
c) sulfurous acid (<u>not</u> "dihydrogen sulfite")   d) copper(II) iodide or cupric iodide
e) nickel nitrate or nickel(II) nitrate   f) lithium sulfide
g) iron(II) phosphate or ferrous phosphate   h) barium carbonate
i) dinitrogen monoxide ("nitrous oxide", or "laughing gas")

2.   a) $NaC_2H_3O_2$ b) NH_4Br c) KCN   d) LiOH
e) $HClO_4(aq)$ f) Al_2O_3 g) $ZnSO_4$ h) SF_6
i) CuO   j) $Ca(HCO_3)_2$

Chapter 4

3.   moles of $CuSO_4 \cdot 5H_2O$ = 8.50 g/249.686 g/mole = 0.0340 mole.
molarity = moles of solute/L of solution = 0.0340 mole/0.250 L = **0.136 M**.

4.   grams of NaOH = (0.18 mole/L)(0.500 L)(39.997 g/mole) = **3.6 g**.

5.   Use the dilution formula: $mL_{conc} \times M_{conc} = mL_{dilute} \times M_{dilute}$

$(mL_{conc})(12\ M) = (1000\ mL)(1.0\ M)$,
$mL_{conc} = (1000\ mL)((1.0\ M)/12\ M = $ **83 mL**.

6.   This is a "two equations, two unknowns" problem.

Let X = L of 1.0 M H_2SO_4 needed,
Let Y = L of 6.0 M H_2SO_4 needed.

Since the total volume needed is 500 mL, or 0.500 L, we know that

X + Y = 0.500 L   (equation 1)

Volume in L $\times$ molarity = moles
The total moles in the 0.500 L of the diluted solution will be

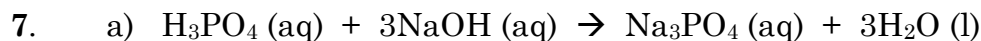0.500 L × 2.5 mole/L = 1.25 moles, so

X(1.0 M) + Y(6.0 M) = 1.25 moles also   (equation 2)

Substitute Y = 0.500 L - X (from equation 1), plug into equation 2, and solve for X:

X(1.0 M) + (0.500 L - X)(6.0 M) = 1.25 mole,

X + 3.0 mole - (6.0 M)(X) = 1.25 mole, or more simply X + 3 - 6X = 1.25,

-5X = -1.75, X = 0.35 L. **350 mL** of 1.0 M H_2SO_4 will be needed.

Y = 0.500 L - 0.350 L = 0.150 L. **150 mL** of 6.0 M H_2SO_4 will be needed.

7.   a) H_3PO_4 (aq) + 3NaOH (aq) → Na_3PO_4 (aq) + $3H_2O$ (l)

b) For a standard acid-base titration problem, we can use our "shortcut" formula:

   mL_{acid} × M_{acid} × n_{H+} = mL_{base} × M_{base} × n_{OH-}

   (25.00 mL)(M_{acid})(3) = (35.92 mL)(0.500 M)(1)

   M_{acid} = **0.239 M**.

c) grams of H_3PO_4 = (0.239 mole/L)(0.02500 L)(97.99518 g/mole) = **0.587 g**.

8.   a) non-electrolyte   b) weak electrolyte   c) strong electrolyte

d) strong electrolyte   e) strong electrolyte   f) weak electrolyte

g) weak electrolyte   h) strong electrolyte   i) non-electrolyte

9.   a) 0.25 M Ca^{2+}(aq), 0.50 M Br^-(aq)

b) 0.30 M Na^+(aq), 0.15 M $SO_4{}^{2-}$(aq)

c) 0.30 M Na^+(aq), 0.10 M $PO_4{}^{3-}$(aq)
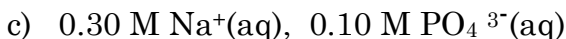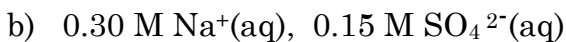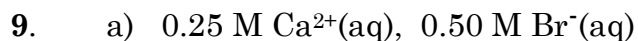
10. **a) Molecular (or "formula") equation:**
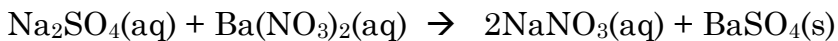
$Na_2SO_4(aq) + Ba(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + BaSO_4(s)$
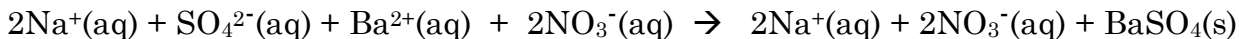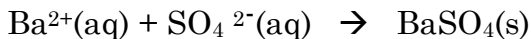
Ionic equation:

$2Na^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2NO_3^-(aq) \rightarrow 2Na^+(aq) + 2NO_3^-(aq) + BaSO_4(s)$

Net ionic equation:

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

b) Molecular equation:

$Pb(NO_3)_2(aq) + Na_2S(aq) \rightarrow 2NaNO_3(aq) + PbS(s)$

Ionic equation:

$Pb^{2+}(aq) + 2NO_3^-(aq) + 2Na^+(aq) + S^{2-}(aq) \rightarrow 2Na^+(aq) + 2NO_3^-(aq) + PbS(s)$

Net ionic equation:

$Pb^{2+}(aq) + S^{2-}(aq) \rightarrow PbS(s)$

c) Molecular equation:

$Na_2CO_3(aq) + 2HCl(aq) \rightarrow H_2CO_3(aq) + 2NaCl(aq)$
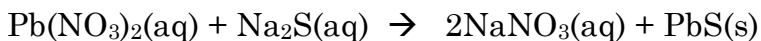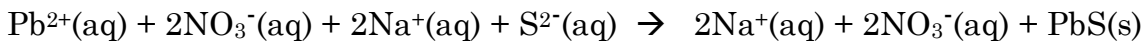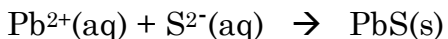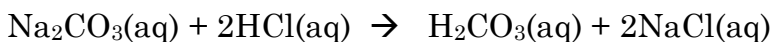
Since $H_2CO_3(aq)$ (carbonic acid) immediately decomposes to $CO_2(g) + H_2O(l)$, we should more correctly write
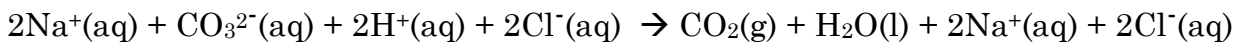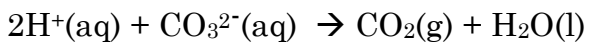
$Na_2CO_3(aq) + 2HCl(aq) \rightarrow CO_2(g) + H_2O(l) + 2NaCl(aq)$

Ionic equation:

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

Net ionic equation:

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

d) Molecular equation:

$AgNO_3(aq) + KBr(aq) \rightarrow KNO_3(aq) + AgBr(s)$

Ionic equation:

$Ag^+(aq) + NO_3^-(aq) + K^+(aq) + Br^-(aq) \rightarrow K^+(aq) + NO_3^-(aq) + AgBr(s)$

Net ionic equation:

$Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$

e) Molecular equation:

$CuSO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + Cu(OH)_2(s)$

Ionic equation:

$Cu^{2+}(aq) + SO_4^{2-}(aq) + 2Na^+(aq) + 2OH^-(aq) \rightarrow 2Na^+(aq) + SO_4^{2-}(aq) + Cu(OH)_2(s)$
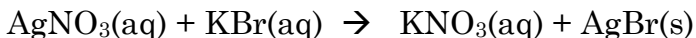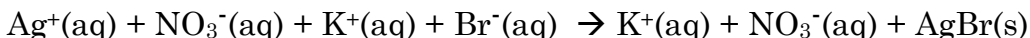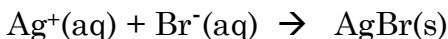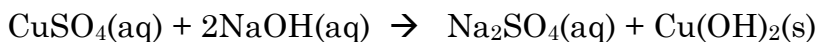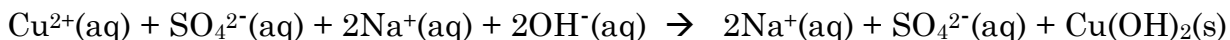
11. a) soluble   b) soluble   c) soluble
          d) soluble   e) insoluble   f) soluble

1) Note that the word "soluble" is <u>not</u> spelled "soluable" !!
2) Use the solubility rules from your textbook or lab manual. You can correctly predict the solubility of many compounds by remembering that <u>Group IA ionic compounds</u> (containing Li$^+$, Na$^+$, K$^+$, etc.), <u>nitrates</u> (containing the NO$_3^-$ ion) and <u>acetates</u> (containing the $C_2H_3O_2^-$ ion) are <u>always</u> soluble in water.

12.   These will be <u>single displacement reactions</u>: A + BX → B + AX. An <u>element</u> A can displace another element (or ion) B from a <u>compound</u> BX <u>only</u> if element A is <u>above</u> element B in the activity series (or, element A must be "more active" than element B). Otherwise, <u>no reaction</u>.

a) Outcomes of reactions:

1) $Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$
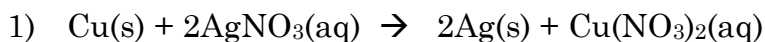
2) $Zn(s) + FeCl_2(aq)$ → $Fe(s) + ZnCl_2(aq)$

3) No reaction (Ni is below Mg in the series)

4) $2Al(s) + 6HCl(aq)$ → $3H_2(g) + 2AlCl_3(aq)$

5) No reaction (Cu is below H in the series)

b) The element which becomes oxidized will have lost one or more electrons, and the element which becomes reduced will have gained one or more electrons. You will notice a pattern in these single displacement reactions: Element A will be oxidized and element B will be reduced in the reaction (A + BX → B + AX).

1) $Cu(s)$ loses two electrons and therefore becomes oxidized to form $Cu^{2+}(aq)$, and each $Ag^+(aq)$ ion gains one electron to become reduced to $Ag(s)$.

2) $Zn(s)$ loses two electrons to become oxidized to $Zn^{2+}(aq)$, and $Fe^{2+}(aq)$ gains two electrons to become reduced to $Fe(s)$.

3) (No reaction, so nothing was oxidized or reduced)

4) $Al(s)$ loses three electrons to become oxidized to $Al^{3+}(aq)$, and each $H^+(aq)$ ion gains one electron to become reduced to form H <u>atoms</u> initially, which immediately combine to form H_2 <u>molecules</u>.

5) (No reaction)
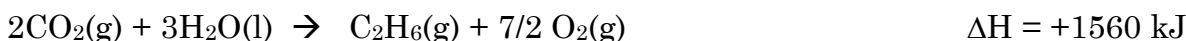
Chapter 6

13. Energy is "conserved" or is not created or destroyed in any process. Energy can change form and be transferred from one substance to another. Mathematically, the first law of thermodynamics can be written as $\Delta E = q + w$.

14. a) $q = +240$ J, $w = -135$ J. $\Delta E = q + w = 240$ J + -135 J = **+105 J**.

b) $q = -225$ J, $w = 0$. $\Delta E = -225$ J + 0 = **-225 J**.

15.   a)   Since ΔH is positive, the reaction is **endothermic** (absorbs heat energy from the surroundings).

   b)   moles of N_2O = 25.0 g/44.0128 g/mole = 0.568 mole.

   (0.568 mole of N_2O)(163.2 kJ/2 moles of N_2O) = **46.3 kJ**.

16.   We need to manipulate the given equations so that they give the desired equation when they are added together. For example, the equation we want has 2C(s) on the left side. Only one of the given equations has C(s) in it (the second one). We need to multiply this equation through by 2 so that we will have <u>2</u>C(s) on the left side, which now makes ΔH for that equation (-394 kJ)(2) = -788 kJ. We will need to <u>reverse</u> the first equation and <u>divide</u> it through by 2 in order to have <u>1</u>C_2H_6(g) on the <u>right</u>. Finally, we must multiply the third equation through by <u>3/2</u> (1.5) in order for the equations to add together to give the desired equation.

   $2C(s) + 2O_2(g) \rightarrow 2CO_2(g)$ ΔH =   -788 kJ

   $2CO_2(g) + 3H_2O(l) \rightarrow C_2H_6(g) + 7/2\ O_2(g)$ ΔH = +1560 kJ

   <u>$3H_2(g) + 3/2\ O_2(g) \rightarrow 3H_2O(l)$</u>   <u>$\Delta H$ =   -858 kJ</u>

   　　　$2C(s) + 3H_2(g) \rightarrow C_2H_6(g)$ **ΔH =   -86 kJ**

17.   From the appendix in our textbook we can obtain the following ΔH_f^o values:

Substance	ΔH_f^o
$NH_3(g)$	-46.19 kJ/mole
$O_2(g)$	0
$NO(g)$	90.37 kJ/mole
$H_2O(\mathbf{g})$	-241.82 kJ/mole

   (Be sure to use ΔH_f^o for water <u>vapor</u> here ($H_2O(\mathbf{g})$), instead of ΔH_f^o for <u>liquid</u> water, $H_2O(\mathbf{l})$, which has a value of <u>-285.83</u> kJ/mole. It is very easy to get these values mixed up<u>!</u>)

$Al(s)$	0
$Fe_2O_3(s)$	-822.16 kJ/mole
$Al_2O_3(s)$	-1669.8 kJ/mole
$Fe(s)$	0

   Remember that <u>ΔH_f^o for elements in their natural states at 25°C and 1 atm pressure</u>

is zero. Such values are often not given in problems because it is expected that you will know this.

a) ΔH_f^o = [(4 moles)(90.37 kJ/mole) + (6 moles)(-241.82 kJ/mole)] -
$$[(4 \text{ moles})(-46.19 \text{ kJ/mole}) + (5 \text{ moles})(0)]$$

= [-1089.44 kJ] - [-184.76 kJ] = **-904.68 kJ**.

b) ΔH_f^o = [(1 mole)(-1669.8 kJ/mole) + (2 moles)(0)] -
$$[(2 \text{ moles})(0) + (1 \text{ mole})(-822.16 \text{ kJ/mole})]$$

= [-1669.8 kJ] - [-822.16 kJ] = **-847.6 kJ**.

18. Heat lost by the aluminum = heat gained by the water + heat gained by the calorimeter. This is a typical coffee cup calorimetry experiment. Note that in this problem, the heat gained by the water and the heat gained by the calorimeter (the coffee cup) are calculated separately. Often, the "calorimeter constant" includes both of these terms (for example, in problem 19a following).

heat lost by the Al = mass of Al $\times$ ΔT of the Al $\times$ specific heat of Al
= (7.23 g)(80.0^{o}C - T_f)(0.90 J/g^{o}C)

heat gained by the water = mass of H_2O $\times$ ΔT of the H_2O $\times$ specific heat of H_2O
= (50.0 g)(T_f - 24.0^{o}C)(4.184 J/g^{o}C)

heat gained by the calorimeter = ΔT of the calorimeter $\times$ the calorimeter constant
(or heat capacity of the calorimeter)
= (T_f - 24.0^{o}C)(12.7 J/^{o}C)

ΔT = value - T_f or T_f - value? In this kind of problem, the ΔT terms should always be <u>positive</u>. Subtract the smaller value from the larger value.

(7.23 g)(80.0^{o}C - T_f)(0.90 J/g^{o}C) = (50.0 g)(T_f - 24.0^{o}C)(4.184 J/g^{o}C) + (T_f - 24.0^{o}C)(12.7 J/^{o}C)

520 .56 - 6.507T_f = 209.2T_f - 5028 + 12.7T_f -304.8

228.407T_f = 5846.16, T_f = **25.6 ^{o}C**.

19.   a) Remember that the heat capacity of the the calorimeter will usually have units of J or kJ per ^{o}C or K (we will use kJ/^{o}C here).

(1.200 g)(26.38 kJ/g) = 31.66 kJ of heat gained by the calorimeter.

heat capacity of the calorimeter = 31.66 kJ/ΔT = 31.66 kJ/(26.10^{o}C - 22.45^{o}C)

= 31.656 kJ/3.65^{o}C = **8.67 kJ/^{o}C**.

b) heat gained by water = (1500 g)(3.65^{o}C)(4.184 J/g^{o}C) = 22907.4 J = 22.91 kJ.

heat gained by the calorimeter + water = 31.66 kJ (calculated in (a) above)

heat gained by the calorimeter alone = 31.66 kJ - 22.91 kJ = 8.75 kJ.

The heat capacity of the calorimeter alone = 8.75 kJ/ΔT = 8.75 kJ/3.65^{o}C

= **2.40 kJ/^{o}C**.

Chapter 7

20.   For these problems, you will be using the equations $\mathbf{\nu = c/\lambda}$, $\mathbf{\lambda = c/\nu}$, and $\mathbf{E\ or\ \Delta E = h\nu = hc/\lambda}$ very often. <u>Memorize them!</u>

ν (nu) = frequency in units of 1/sec or sec^{-1} or Hz
λ (lambda) = wavelengh which needs to be converted to <u>meters</u> if <u>energy</u> is being
calculated
c = speed of light in vacuum = 3.00 X 10^{8} m/sec
h = Planck's constant = 6.63 X 10^{-34} J sec
E = energy in J

a) λ = c/ν = (3.00 X 10^{8} m/sec)/(820 X 10^{3} sec^{-1}) = **366 m**   (AM station)

λ = c/ν = (3.00 X 10^{8} m/sec)/(89.7 X 10^{6} sec^{-1}) = **3.34 m**   (FM station)

b) $E = h\nu = (6.63 \times 10^{-34} \text{ J sec})(820 \times 10^3 \text{ sec}^{-1}) = \mathbf{5.44 \times 10^{-28} \text{ J}}$ (AM station)

   $E = h\nu = (6.63 \times 10^{-34} \text{ J sec})(89.7 \times 10^6 \text{ sec}^{-1}) = \mathbf{5.95 \times 10^{-26} \text{ J}}$ (FM station)

21. $\nu = c/\lambda$, $E = h\nu = hc/\lambda = (6.63 \times 10^{-34} \text{ J sec})(3.00 \times 10^8 \text{ m/sec})/(785 \times 10^{-9} \text{ m})$,

   $E = 2.53 \times 10^{-19}$ J per photon.

   $(31 \text{ J})/(2.53 \times 10^{-19} \text{ J per photon}) = \mathbf{1.2 \times 10^{20} \text{ photons}}$.

22. a) emitted   b) absorbed   c) absorbed   d) absorbed

23. $\Delta E = R_H (1/n_i^2 - 1/n_f^2)$, $R_H = 2.18 \times 10^{-18}$ J

   a) $n_i = 1$, $n_f = 5$. $\Delta E = (2.18 \times 10^{-18} \text{ J})(1/1^2 - 1/5^2) = \mathbf{2.09 \times 10^{-18} \text{ J}}$.

     $\Delta E = h\nu$, $\nu = \Delta E/h = (2.09 \times 10^{-18} \text{ J})/(6.63 \times 10^{-34} \text{ J sec}) = \mathbf{3.16 \times 10^{15} \text{ sec}^{-1}}$.

     $\lambda = c/\nu = (3.00 \times 10^8 \text{ m/sec})/(3.16 \times 10^{15} \text{ sec}^{-1}) = \mathbf{9.50 \times 10^{-8} \text{ m}}$.

   b) $\Delta E = (2.18 \times 10^{-18} \text{ J})(1/1^2 - 1/\text{infinity}^2) = \mathbf{2.18 \times 10^{-18} \text{ J}}$. (This is the
     <u>ionization energy</u> of the hydrogen atom, or the energy required to completely
     remove the electron from the hydrogen atom.)

     $\nu = \mathbf{3.29 \times 10^{15} \text{ sec}^{-1}}$, $\lambda = \mathbf{9.12 \times 10^{-8} \text{ m}}$.

24. We will use deBroglie's equation $\lambda = h/mv$. Remember that the joule unit in
   Planck's constant h $(6.63 \times 10^{-34}$ J sec) is in kg m²/sec², so we need to convert mass
   to kilograms, length to meters, and time to seconds. We'll also wait with rounding
   until we get to the final answer, which should be rounded to two significant figures.

   $m = (5.0 \text{ oz})(28.3 \text{ g/oz})(1 \text{ kg}/1000 \text{ g}) = 0.1415$ kg

   $v = (89 \text{ mi/hr})(1.6093 \times 10^3 \text{ m/mi})(1 \text{ hr}/60 \text{ min})(1 \text{ min}/60 \text{ sec}) = 39.785$ m/sec.

   $\lambda = (6.63 \times 10^{-34} \text{ kg m}^2 \text{ sec}^{-2} \cdot \text{sec})/(0.1415 \text{ kg})(39.785 \text{ m sec}^{-1}) = \mathbf{1.2 \times 10^{-34} \text{ m}}$.

25. The symbol **ψ** (psi, pronounced like "sigh") represents mathematical functions called *wave functions* or *wave equations* in quantum mechanics. All wave functions are solutions to a general equation known as the *Schrödinger equation*. The wave function gives us information about the distribution of the electrons in the space about the nucleus.

The *square* of ψ, ψ X ψ or **ψ²**, is related to the *probability* of finding the electron in a given region about the nucleus. To illustrate what this means, I will borrow a part of James Plambeck's excellent chemistry notes (see http://www.chem.ualberta.ca/~plambeck/che/p101/p01233.htm for his complete discussion):

"A plot of the wave function ψ^2 against radial distance from the nucleus is shown for hydrogen s orbitals in the figure below. This plot, which represents the relative probability of finding the electron per unit volume, is difficult to visualize physically because there is much less volume close to the nucleus than there is further out. It is easier to visualize where the electron may be by plotting the <u>radial probability density</u> $4\pi\, r^2\psi^2$, rather than ψ^2, against the distance from the nucleus as shown on the right in the figure."
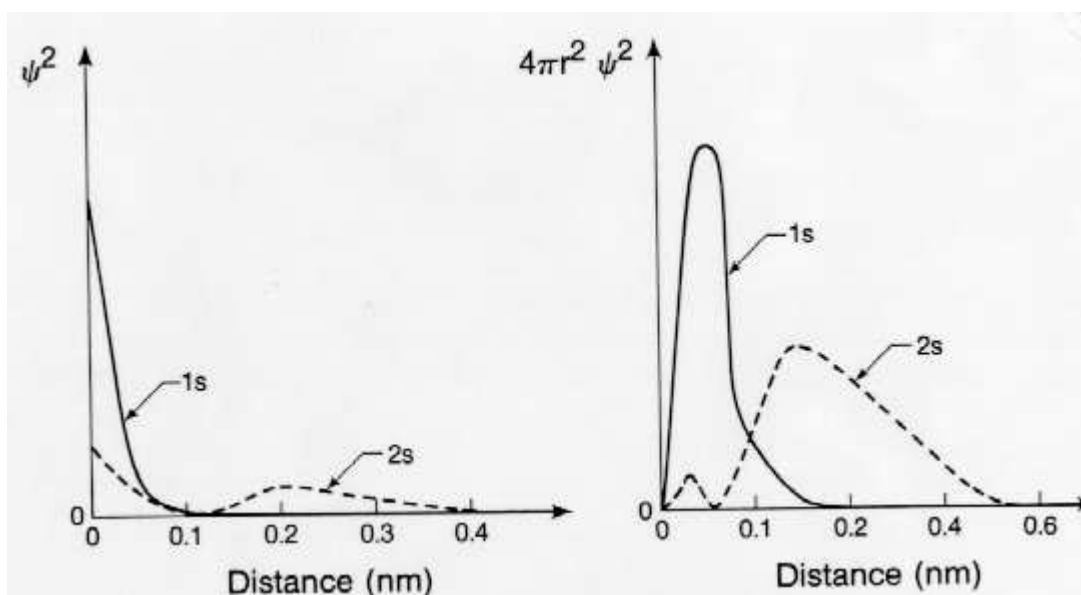
We interpret radial probability density as the probability of finding the electron within a very thin spherical region at different distances (radii) from the nucleus. This is a more realistic representation which also eliminates the confusion seen in the ψ^2 plot in which the probability of finding the electron is greatest at r=0, the exact *center* of the atom, which doesn't make sense. Unfortunately most chemistry textbooks (including ours) do not address this subject sufficiently.

26. This is a good place to contrast the picture of the atom presented by the Bohr model with the results of the quantum mechanical model. The Bohr model of the hydrogen atom has the electron orbiting the nucleus in a well-defined path just as the earth orbits the sun:
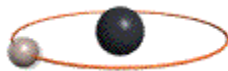
The quantum mechanical model gives *wave functions* which describe the region about the nucleus where the electron will be found in terms of *probability*; we can never experimentally determine exactly where an electron is at any given instant.

Shown below are space-filling representations of the 1s, 2p, and 3d orbitals. There is about a 95% probability that an electron in one of these orbitals will be within these pictured regions. They are not to scale. Keep in mind the cross sections shown in problem 25 for the 1s and 2s orbitals (the $4\pi\, r^2\psi^2$ *vs* r plots); the p and d orbitals have similar patterns.
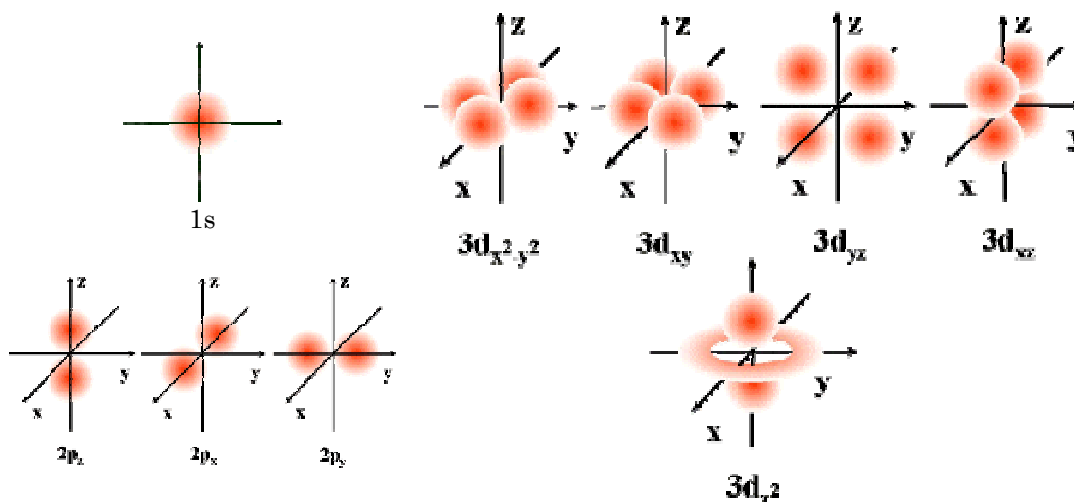
If you are interested, you can find illustrations of the *f*-orbitals at, for example, http://antoine.frostburg.edu/chem/senese/101/electrons/faq/f-orbital-shapes.shtml.

What do the letters s, p, d, and f in the orbital names stand for?   Fred Senese (http://antoine.frostburg.edu/chem/senese/101/history/faq/why-spdf.shtml) writes:

"You might expect that the 's' stands for 'spherical' and 'p' stands for 'polar' because these imply the shapes of the *s* and *p* orbitals, but unfortunately, the letter designations have nothing to do with the orbital shapes. Spectroscopists associated transitions involving energy levels with different values with different groups of lines in the line spectra of the alkali metals. The line groups were called *sharp*, *principal*, *diffuse*, and *fundamental*. When the angular momentum quantum number was used to describe and explain these groups of lines, s became an abbreviation for = 0, p meant = 1, d meant = 2, and f meant = 3. For consistency, higher values of the angular momentum quantum numbers are designated alphabetically (g means = 4, h means = 5, and so on)."

Quantum Numbers

The first three quantum numbers n, ℓ, and m_ℓ come from mathematical functions which are solutions to the *Schrödinger equation*. These equations are very complex. A portion is shown below just to give you an idea:

$$u_{n\ell}(\rho) = \left(\frac{2Z}{n\,a_0}\right)^{1/2} \sqrt{\frac{(n-\ell-1)!}{2n\,((n+1)!)^3}}\; e^{-Z\rho/n} \left(\frac{2Z\rho}{n}\right)^{\ell+1} L_{n-\ell-1}^{(2\ell+1)}\left(\frac{2Z\rho}{n}\right)$$

Wow! But from this, we can at least get an idea of why the quantum numbers are allowed to have only certain values. Look at the first term in this equation; if n = 0 we will have a zero denominator giving an undefined result, therefore, n cannot be zero. Looking at the first and second terms, we can see that when n is negative or if n = ℓ, we will be taking the square root of a negative number, also a problem. While this material is beyond the scope of our course, if you are interested, you can find more details about these wave functions at http://www.nat.vu.nl/~wimu/Schrod.html.

The fourth quantum number, m_s, is a property of the electron itself and is not part of the wave function. Following is a summary of the four quantum numbers and their allowed values. Memorize!

The principle quantum number, n, describes the general energy level; as n increases the electron will be in higher energy levels farther away from the nucleus.

> n = 1, 2, 3, 4, 5, ... (positive integers up to infinity)

The secondary or azimuthal quantum number, ℓ, describes the type of orbital the electron is in. ℓ = 0 = s-orbital, ℓ = 1 = p-orbital, ℓ = 2 = d-orbital, ℓ = 3 = f-orbital, etc.

> ℓ = 0, 1, 2, 3, 4, ... (up to n-1)

The magnetic quantum number, m_ℓ (pronounced "m sub el") describes the orientation of the orbital.

> m_ℓ = ... -4, -3, -2, -1, 0, +1, +2, +3, +4, ... (integer values from $-\ell$ to $+\ell$)

Finally, the spin quantum number, m_s (pronounced "m sub s") gives the "spin" of the electron itself. An electron behaves like (but is probably not really) a spinning charged solid object; such an object generates a magnetic field whose orientation depends on the direction of spin.

> m_s = +½ or -½

27. A *node* is a region where the probability of finding the electron (ψ^2 or $4\pi\, r^2\psi^2$) is <u>zero</u>. A good illustration of the nodes in s-orbitals can be found at http://cw.booksites.net/~bookbind/pubbooks/blb/chapter6/medialib/blb0606.html (link out-of-date, quoted below; see also http://wps.prenhall.com/wps/media/objects/3311/3390683/blb0606.html):
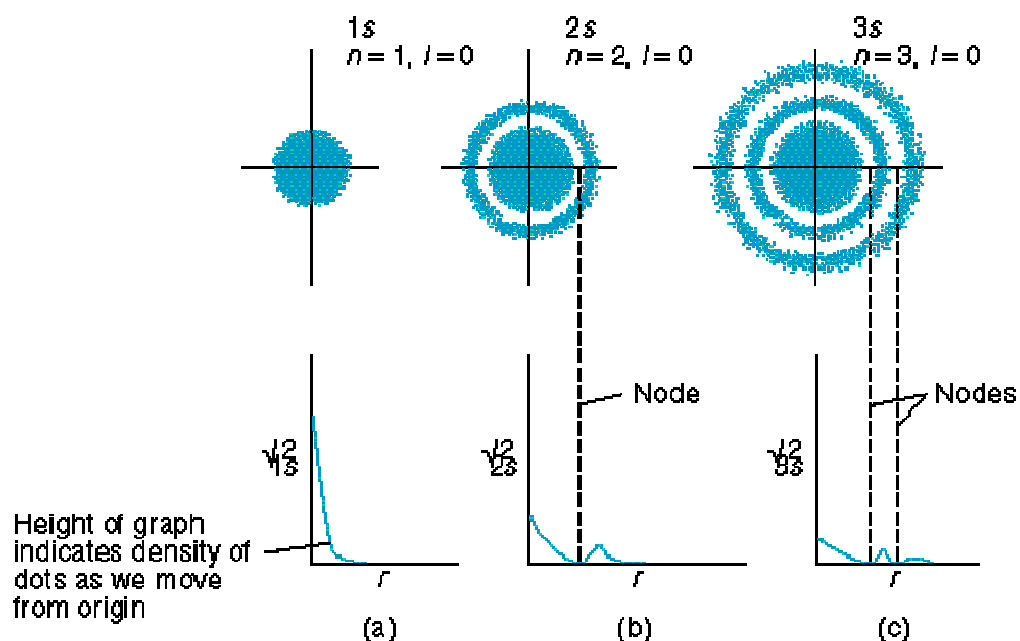
"The lower part of the figure shows how the electron density, represented by ψ^2, varies as a function of distance r from the nucleus. In the 2s and 3s orbitals the electron-density function drops to zero at certaindistances from the nucleus. The spherical surfaces around the nucleus at which ψ^2 is zero are called nodes.

"The lowest-energy orbital, the 1s orbital, is spherical, as shown in (the figure). Figures of this type, showing electron density, are one of the several ways we use to help us visualize orbitals. This figure indicates that the probability of finding the electron around the nucleus decreases as we move away from the nucleus in any direction. When the probability function, ψ^2, for the 1s orbital is graphed as a function of the distance from the nucleus, r, it rapidly approaches zero, as shown in (a). This effect indicates that the electron, which is drawn toward the nucleus by electrostatic attraction, is unlikely to be found very far from the nucleus.

"If we similarly consider the 2s and 3s orbitals of hydrogen, we find that they are also spherically symmetric. Indeed, all s orbitals are spherically symmetric. The manner in which the probability function, ψ^2, varies with r for the 2s and 3s orbitals is shown in (b) and (c). Notice that for the 2s orbital, ψ^2 goes to zero and then increases again in value before finally approaching zero at a larger value of r. The intermediate regions where ψ^2 goes to zero are called nodes. The number of nodes increases with increasing value for the principal quantum number, n. The 3s orbital possesses two nodes, as illustrated in (c). Notice also that as n increases, the electron is more and more likely to be located farther from the nucleus. That is, the size of the orbital increases as n increases."

28. You will find definitions of these terms are in your textbook. See if you can define them to yourself verbally. Practice!

29. You can write electron configurations following the scheme we started with in class:

$$1s \ 2s \ 3s \ 4s \ 5s \ 6s \ 7s \ 8s \ ...$$
$$2p \ 3p \ 4p \ 5p \ 6p \ 7p \ 8p \ ...$$
$$3d \ 4d \ 5d \ 6d \ 7d \ 8d \ ...$$
$$4f \ 5f \ 6f \ 7f \ 8f \ ...$$
$$5g \ 6g \ 7g \ 8g \ ...$$
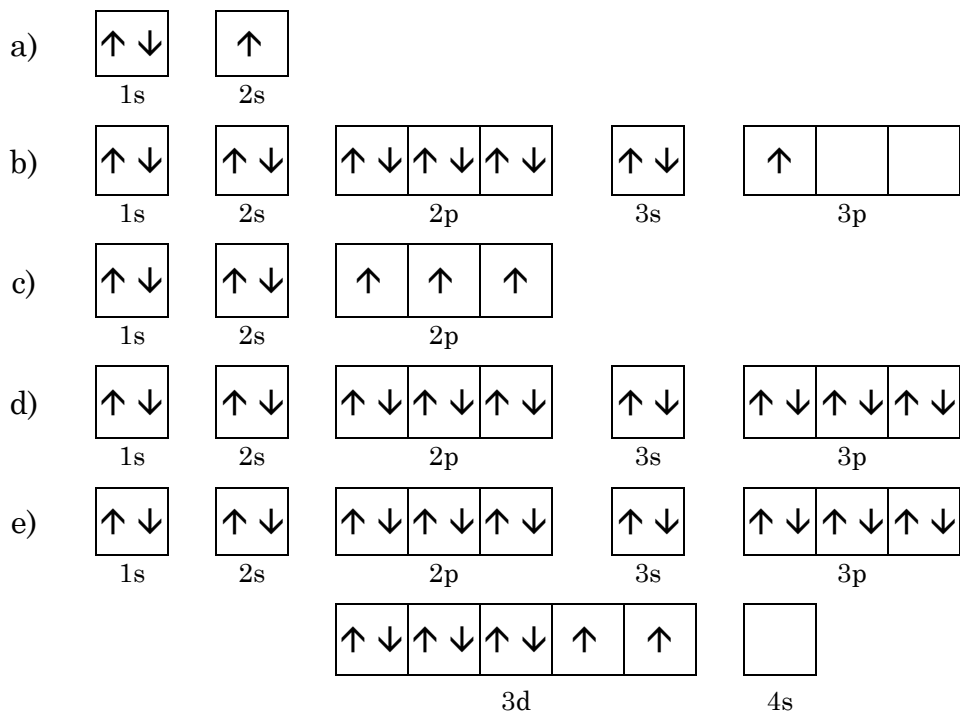
However, it is much better practice to use the periodic table as your guide. Once again, practice makes perfect!

a) $_6C$ $1s^2 \ 2s^2 \ 2p^2$ or   [He] $2s^2 \ 2p^2$

b) $_{15}P$ $1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^3$ or   [Ne] $3s^2 \ 3p^3$

c) $_{10}Ne$ $1s^2 \ 2s^2 \ 2p^6$ or   [Ne]

d) $_{19}K$ $1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^1$ or   [Ar] $4s^1$

e) $_{11}Na^+$ $1s^2 \ 2s^2 \ 2p^6$ or   [Ne]

f) $_{17}Cl^-$ $1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6$ or   [Ar]

g) $_{26}Fe^{3+}$ $1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^0 \ 3d^5$ or   [Ar] $3d^5 \ 4s^0$
                                    (<u>not</u> $4s^2 \ 3d^3$)

h) $_{92}U$ $1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^2 \ 3d^{10}$
             $4p^6 \ 5s^2 \ 4d^{10} \ 5p^6 \ 6s^2 \ 5d^1 \ 4f^{14}$
             $5d^9 \ 6p^6 \ 7s^2 \ 6d^1 \ 5f^3$ or   [Rn] $7s^2 \ 6d^1 \ 5f^3$
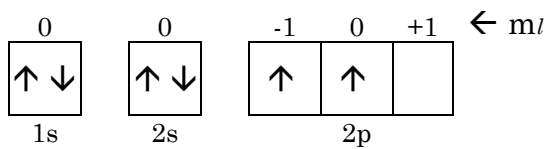
           (using the periodic table in our textbook)

30. Orbital ("box") diagrams

a) = Li   b) = Al   c) = N   d) = S^{2-} e) = Ni^{2+}

a) [↑↓] [↑]
   1s   2s

b) [↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓] [↑| |]
   1s   2s   2p   3s   3p

c) [↑↓] [↑↓] [↑|↑|↑]
   1s   2s   2p

d) [↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓] [↑↓|↑↓|↑↓]
   1s   2s   2p   3s   3p

e) [↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓] [↑↓|↑↓|↑↓]
   1s   2s   2p   3s   3p

        [↑↓|↑↓|↑↓|↑|↑] [ ]
            3d   4s

31. Orbital diagram for carbon, $1s^2\ 2s^2\ 2p^2$

  0   0   -1 0 +1 ← m_l
[↑↓] [↑↓] [↑|↑|]
 1s   2s   2p

Electron #	n	ℓ	m_ℓ	m_s
1	1	0	0	+ ½
2	1	0	0	- ½
3	2	0	0	+ ½
4	2	0	0	- ½
5	2	1	-1	+ ½
6	2	1	0	+ ½

Value of ℓ	Type of Orbital
0	s
1	p
2	d
3	f
4	g
...	...

32. a) <u>excited state</u> (The first two electrons have the same four quantum numbers and therefore violates the Pauli exlusion principle.)

b) <u>ground state</u> (You can put electrons into <u>degenerate</u> orbitals in any order, but I usually put them in going from left to right for consistency.)

c) <u>excited state</u> (The lower energy levels should be filled first, then the next highest, and so on. This is called the <u>aufbau principle</u> (German, "filling up").)

d) <u>excited state</u> (<u>Hund's rule</u> says that when putting electrons into degenerate orbitals, the most stable configuration results when these electrons are first put in <u>singly</u> (one electron per orbital) with their <u>spins parallel</u> (all $+\frac{1}{2}$ or all $-\frac{1}{2}$).)

33. a) allowed   b) not allowed (m_ℓ can only be zero if $\ell = 0$)

c) allowed   d) allowed

34. The aluminum ion, Al^{3+}, results when an aluminum atom loses its three <u>outermost</u>, or <u>valence</u>, electrons:

$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1\ -3e\ \rightarrow\ \ \ 1s^2\ 2s^2\ 2p^6$$

or $[Ne]\ \ \ \ 3s^2\ 3p^1\ -3e\ \rightarrow\ \ \ [Ne]\ 3s^0\ 3p^0\ =\ [Ne]$
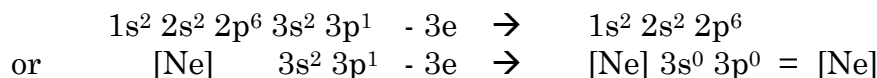
The Al^{3+} ion has the same electron configuration as the noble gas neon, a very <u>stable</u> atom. If we were to try to remove a <u>fourth</u> electron, it would have to be one of the <u>inner</u>, or <u>core</u>, electrons, much more difficult to remove. Because this inner core is so stable, core electrons are not normally lost in chemical reactions.

Similarly, an oxygen atom can achieve a noble gas configuration by <u>gaining</u> two electrons:

$$1s^2\ 2s^2\ 2p^4\ \ +2e\ \rightarrow\ \ \ 1s^2\ 2s^2\ 2p^6$$

or $[He]\ 2s^2\ 2p^4\ +2e\ \rightarrow\ \ \ [He]\ 2s^2\ 2p^6\ =\ [Ne]$
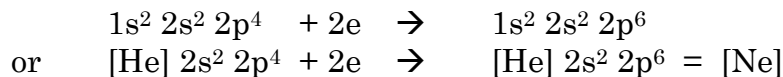
Thus oxygen tends to <u>gain</u> two electrons in its reactions, forming the oxide ion (O^{2-}) which occurs in <u>ionic</u> compounds of oxygen, or through <u>sharing</u> electrons with another atom, usually another nonmetal in this case, forming <u>molecular</u> compounds. You might ask whether oxygen ever loses, for example, its six valence electrons to form a +6 ion with a [He] configuration. The answer is, it doesn't; it becomes more and more difficult (requires more energy) to remove electrons from ions with high positive charges ($O \rightarrow\ O^+ \rightarrow O^{2+} \rightarrow O^{3+} \rightarrow O^{4+} \rightarrow O^{5+} \rightarrow O^{6+}$). As a rule, elements tend to gain or lose electrons to achieve the configuration of the noble gas <u>nearest</u> to it. A general consequence of this is that <u>metals tend to lose electrons in reactions and nonmetals tend to gain electrons in reactions</u>.