Ch 3 and 4 – Advanced Calculations/Problems **Name:** Key
Honors Chemistry

Formulas and constants:

Avogadro's number = 6.02×10^{23} particles/mol	$c=\lambda\nu$
$c= 3.00 \times 10^{8}$ m/s	$E=h\nu$
$h = 6.626 \times 10^{-34}$ J·s	$E_n = \frac{-2.178 \times 10^{-18}\text{ J}}{n^2}$
1 nm = 1×10^{-9}m	
1MHz = 1×10^{6}Hz	$\Delta E = E_f - E_i$

1. Answer the following questions about glucose, $C_6H_{12}O_6$.

   a. How many molecules of glucose are in 85.0 grams of glucose?

   $85.0\text{ g} \times \frac{1\text{ mol}}{180\text{ g}} \times \frac{6.02\times10^{23}\text{ molecules}}{1\text{ mol}} = 2.84\times10^{23}\text{ molecules}$

   b. How many atoms of carbon are in 85.0 grams of glucose?

   $2.84\times10^{23}\text{ molecules } C_6H_{12}O_6 \times \frac{6\text{ atoms C}}{1\text{ molecule}} = 1.70\times10^{24}\text{ atoms C}$

Fix → c. What is the mass (in kg) of 1.20×10^{26} ~~moles~~ molecules of glucose?

   $1.20\times10^{26}\text{ molecules} \times \frac{1\text{ mol}}{6.02\times10^{23}\text{ molecules}} \times \frac{180\text{ g}}{1\text{ mol}} \times \frac{1\text{ kg}}{1000\text{ g}} = 35.9\text{ kg}$

   $1.20\times10^{26}\text{ moles} \times \frac{180.0\text{ g}}{1\text{ mol}} \times \frac{1\text{ kg}}{1000\text{ g}} = 2.16\times10^{25}\text{ kg}$

2. Magnetic Resonance Imaging (MRI) is a powerful diagnostic tool used in medicine. The imagers used in hospitals operate with a wavelength of 7.50×10^{8} nm. Calculate $= 0.750\text{ m}$

   a. The frequency in MHz.

   $c=\lambda\nu$

   $\nu = \frac{c}{\lambda}$

   $\nu = \frac{3.00\times10^{8}\text{ m/s}}{0.750\text{ m}} = 4.00\times10^{8}\text{ Hz} \times \frac{1\text{ MHz}}{1\times10^{6}\text{ Hz}} = 400.\text{ MHz}$

   b. The energy in J/photon

   $E = h\nu$

   $= (6.626\times10^{-34}\text{ J·s})(4.00\times10^{8}\text{ Hz}) = 2.65\times10^{-25}\text{ J/photon}$

   c. The energy in kJ/mol of photons

   $\frac{2.65\times10^{-25}\text{ J}}{1\text{ photon}} \times \frac{6.02\times10^{23}\text{ photons}}{1\text{ mol photons}} \times \frac{1\text{ kJ}}{1000\text{ J}} = 1.60\times10^{-4}\text{ kJ/mol of photons}$

3. Consider the following electron transitions in a hydrogen atom:

   i. n=3 to n=1
   ii. n=2 to n=3
   iii. n=4 to n=3
   iv. n=3 to n=5

a. Which of the transitions involves the absorption of energy by the atom?

ii and iv (increase in energy levels)

b. In which is energy emitted from the atom? (There are two answers)

i and iii

c. Calculate the amount of energy emitted in each of the transitions listed in "b".

(i) $E_3 = \frac{-2.178\times10^{-18} J}{9} = -2.42\times10^{-19} J$ $\quad E_1 = -2.178\times10^{-18} J$

$\Delta E = E_1 - E_3 = -1.94\times10^{-18} J$

(iii) $E_4 = \frac{-2.178\times10^{-18} J}{16} = 1.36\times10^{-19} J$ $\quad E_3 =$ (see above)

$\Delta E = E_3 - E_4 = -1.06\times10^{-19} J$

4. Given the following sets of electron quantum numbers, indicate those that could not occur and explain your answers.

   a. 1, 0, 0, -1/2
   b. 2, 2, 1, +1/2
   c. 3, 2, -2, +1/2
   d. 4, 0, 2, -1/2

a) ↑↓ 1s

b) refers to 2d sublevel which does not exist

$n = 2$, $l = 2$ (Not possible)

$l = 0 \ldots (n-1)$

c) ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ 3d

d) refers to an orbital that does not exist in an "s" sublevel

$m_l = -l \ldots 0 \ldots +l$ ($\therefore$ if $l = 0$ then $m_l = 0$)