Name: _____   Student Number: _____

Chemistry 1000 Practice Final Exam C
Based on Fall 2014 Test (Content Updated to Fall 2017 Curriculum)

INSTRUCTIONS
1) Read the exam carefully before beginning. There are 21 questions on pages 2 to 13 followed by 2 pages of "Data Sheet" (including periodic table) and a blank page for any rough work. **Please ensure that you have a complete exam. If not, let an invigilator know immediately**. All pages must be submitted at the end of the exam.
2) If your work is not legible, it will be given a mark of zero.
3) Marks will be deducted for incorrect information added to an otherwise correct answer.
4) You may use a calculator.
5) Show your work for all calculations. Answers without supporting calculations will not be given full credit.
6) Marks will be deducted for improper use of significant figures and for numerical answers with incorrect/missing units.
7) **Do not open the exam until you are told to begin. Beginning prematurely will result in removal of your exam paper and a mark of 0.**
8) You have **3 hours** to complete this exam. Nobody may leave the exam room during the first hour or the last 15 minutes of the exam.

Q	Mark	Q	Mark
1	/ 17	12	/ 4
2	/ 8	13	/ 5
3	/ 8	14	/ 8
4	/ 3	15	/ 2
5	/ 6	16	/ 4
6	/ 9	17	/ 8
7	/ 2	18	/ 11
8	/ 7	19	/ 10
9	/ 6	20	/ 6
10	/ 10	21	/ 6
11	/ 14	22	/ 1

Total	/ 145

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1.   Fill in each blank with the word or short phrase that best completes the sentence. **[17 marks]**

(a)   The ionic radius of Ca^{2+} is <u>smaller</u> than the atomic radius of Ca.

(b)   The atomic radius of neon is <u>smaller</u> than the atomic radius of nitrogen.

(c)   The element with the largest first ionization energy is <u>helium</u>.

(d)   Two elements for which a diagonal relationship (to each other) is observed are <u>lithium</u> and <u>magnesium</u> *or* <u>beryllium</u> and <u>aluminium</u> *or* <u>boron</u> and <u>silicon</u>.

(e)   An atom that has no lone pairs and is bonded to six other atoms has <u>octahedral</u> molecular geometry according to VSEPR theory.

(f)   An atom that has one lone pair and is bonded to three other atoms has <u>trigonal pyramidal</u> molecular geometry according to VSEPR theory.

(g)   Energy is <u>absorbed</u> when a bond is broken.

(h)   One class of nuclear reaction that is not normally spontaneous is <u>fission</u> *or* <u>transmutation</u> *(also called bombardment)*.

(i)   The type of ionizing radiation with the lowest penetrating power is <u>alpha</u>.

(j)   The SI unit for measuring effective biological dose (or equivalent dose) is the <u>sievert</u>.

(k)   The ore from which aluminium is commonly obtained is called <u>bauxite</u>.

(l)   A Lewis base is defined as <u>an electron pair donor</u>.

(m) $HClO_4$ is a <u>stronger</u> Brønsted acid than $HClO_3$.

(n) Ti^{3+} is a <u>weaker</u> Lewis acid than Ti^{4+}.

(o)   If two gases are kept at the same temperature, the particles of the gas with the higher molar mass will have a <u>slower/lower</u> average speed.

(p)   The only intermolecular forces active in a pure sample of a neutral nonpolar substance are <u>London forces</u> *or* <u>London dispersion forces</u> *or* <u>induced dipole-induced dipole forces</u>.

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2.   Complete the following table: **[8 marks]**

Z	Name of Element	Charge of Most Commonly Formed Ion
4	beryllium	+2
21	scandium	+3
28	nickel	+2
34	selenium	-2
35	bromine	-1

Misspelled element names only get partial credit. If the misspelling is severe enough that it no longer matches the name (e.g. "selium" instead of "selenium"), no credit is given.

3.   Complete the following table: **[8 marks]**

Symbol	Z	N	Stable? (yes/no)	Predicted Type of Decay*
$^{186}_{83}\text{Bi}$	83	103	no	alpha (α)
$^{60}_{22}\text{Ti}$	22	38	no	beta (β^-)

*If you have indicated that an isotope is stable, do not fill in the box for its "predicted type of decay".

4.   Complete the following table.   **[3 marks]**

Name	Formula
copper(II) chloride	$CuCl_2$
chromium(III) nitrite	$Cr(NO_2)_3$
aluminium oxide	Al_2O_3

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5.   Write the electron configuration for each of the following atoms/ions.
      Please use a noble gas abbreviation.   **[6 marks]**

(a)   Zn $[Ar]\ 4s^2\ 3d^{10}$

(b) S^{2-} $[Ne]\ 3s^2\ 3p^6$

(c) Mn^{2+} $[Ar]\ 3d^5$

6.   **[9 marks]**

(a)   Write the electron configuration of a ground-state phosphorus atom.
       Do not use a noble gas abbreviation.   *[2 marks]*
              $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

(b)   Draw an orbital occupancy diagram (also called "orbital box diagram") for the **valence
       electrons** of a ground-state phosphorus atom. *Label each subshell.*   *[3 marks]*

3s   **3p**
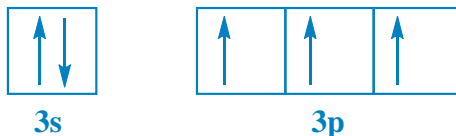
(c)   Use the table below to list a set of quantum numbers describing the **valence electrons** in a
       ground-state phosphorus atom.   *[4 marks]*
       Use as many rows as necessary; the correct answer may include one or more empty rows.

Electron	n	l	m_l	m_s
1 (3s)	3	0	0	+½
2 (3s)	3	0	0	-½
3 (3p)	3	1	-1	+½ **
4 (3p)	3	1	0	+½ **
5 (3p)	3	1	+1	+½ **
6				
7				
8				

  * All three 3p electrons must have different values for m_l. Do not list the range of acceptable
     values! That doesn't tell us that you know they must be different!
 ** Alternatively, all three 3p electrons can have $m_s = -½$. They must all have the same m_s.

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7.   **[2 marks]**

(a)   State the Pauli exclusion principle.

Two electrons in the same system (atom or molecule) cannot have the same quantum state (set of quantum numbers).

(b)   In atoms, what is the consequence of the Pauli exclusion principle?

Since every electron must have a unique set of quantum numbers and there are only two valid values for m_s, there can be a maximum of two electrons in an atomic orbital. (Every electron in a particular orbital must have the same values for n, l, and m_l.)

8.   **[7 marks]**

(a)   Write a balanced chemical equation for the reaction of nitrogen with hydrogen in the presence of a suitable catalyst (Haber process or Haber-Bosch process).   *[2 marks]*
      You must include states of matter.

$$N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g)$$

(b)   In addition to the catalyst, what reaction conditions must be used to generate the product at a reasonable rate and in good yield? Explain briefly why these conditions are necessary.
                                                                      [3 marks]

Heat and high pressure are both necessary.

Heat is necessary because the bonds that must be broken for this reaction to proceed are very strong. The H-H bond is strong because the hydrogen atoms are very small (even relative to other atoms) so the electrons in the bond are close to both nuclei. The N-N bond is strong because it is a triple bond.

High pressure is necessary because it promotes the collisions necessary for the reaction to proceed. (Molecules cannot react without a collision.) At higher pressures, the "concentration" of gas molecules is higher so the probability of one striking the catalyst is higher.

(c)   Agrium operates a large Haber-Bosch plant near Carseland, Alberta. The hydrogen required for this process is obtained by steam reforming of natural gas, followed by the water gas-shift reaction. Write balanced equations for both of these reactions.   *[2 marks]*

$$CH_4(g) + H_2O(g) \rightarrow CO(g) + 3H_2(g)$$

$$CO(g) + H_2O(g) \rightarrow CO_2(g) + H_2(g)$$
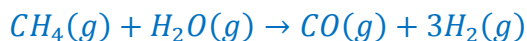
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9. The following reactions studied in class are important to the production and setting of lime mortar. Write a balanced chemical equation for each of these reactions.   **[6 marks]**
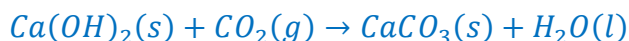   *You must include states of matter.*

(a) Calcium carbonate is heated to give calcium oxide.

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

(b) Calcium oxide is hydrated to give calcium hydroxide.

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$$

(c) Calcium hydroxide absorbs carbon dioxide to regenerate calcium carbonate.

$$Ca(OH)_2(s) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

10. For each of the following substances, *either* write a balanced chemical equation for its reaction with water at room temperature *or* write "no reaction" if it does not react with water at room temperature. *You must include states of matter.*   **[10 marks]**
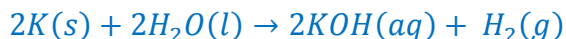
(a) aluminium metal

no reaction

(b) calcium metal

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$$

(c) potassium metal

$$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$$

(d) SO$_3$ gas

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

(e) N$_2$ gas

no reaction

(f) HCl gas

$$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

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11.   **[4 marks]**

(a) Write a balanced chemical equation for the electrolysis of **liquid** sodium chloride in a Down's cell. *You must include states of matter.*   *[2 marks]*

$$2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$$

(b) Write a balanced chemical equation for the reaction that will occur if the products of this electrolysis are not kept separate from each other. *You must include states of matter.[2 marks]*

$$2Na(l) + Cl_2(g) \rightarrow 2NaCl(l)$$

12. An alkaline earth metal (1.00 g) is reacted with chlorine gas to produce the resulting chloride salt (1.81 g). A solution of the salt in water gives a red flame test. Identify the alkaline earth metal **and** give the chemical formula for the chloride salt.   **[4 marks]**

You must show your work and/or explain your logic. You may find it easiest to refer to the metal as M until you identify it. For full credit, your answer must address the masses.

symbol for alkaline earth metal = ___Sr____   formula for chloride salt = __SrCl$_2$__

All alkaline earth metals make +2 cations, so the chloride salt must have the formula MCl$_2$.

The red colour of the flame test suggests that the metal should either be Ca or Sr.

One approach:

1.00 g M reacts with Cl$_2$ to make 1.81 g MCl$_2$, so $m_{Cl_2} = 0.81$g (law of conservation of mass).

Therefore $n_{Cl_2} = 0.81 \ g \ \times \frac{1 \ mol}{70.9054 \ g} = 0.011 \ mol$

Therefore $n_M = 0.011 \ mol \ Cl_2 \ \times \frac{1 \ mol \ M}{1 \ mol \ Cl_2} = 0.011 \ mol \ M$

Therefore $M_M = \frac{1.00 \ g}{0.011 \ mol} = 88 \frac{g}{mol}$ 2 sig. fig.

Therefore M is Sr.

|	M(s)	+	Cl$_2$(g)	→	MCl$_2$(s)
M	???		70.9054 g/mol		(??? + 70.9054) g/mol
m$_{initial}$	1.00 g		0.81 g		1.81 g
n$_{initial}$	0.011 mol		0.011 mol		0 mol
M	88 g/mol				158 g/mol

Other common approaches used by students:

- *Use the ratio of the mass of M and the mass of MCl$_2$ to calculate what the ratio of their molar masses must be:* $\left(\frac{1.00 \ g}{1.81 \ g}\right) = \left(\frac{x}{x+70.9054}\right)$ *then solve for x. This type of approach is called "proportional reasoning".*
- *Trial and error. Calculate n_M and n_{MCl_2} for M = Ca and for M = Sr (or for all cases of M is an alkaline earth metal) and see which metal gives the same value for both*

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13. The hydride ion (H^-) is a strong base. It reacts with water to give the hydroxide ion and hydrogen gas. Write a reaction equation for this process, drawing Lewis diagrams for all reactants and products. Also, use curly arrows to show the movement of electrons in this reaction. **[5 marks]**
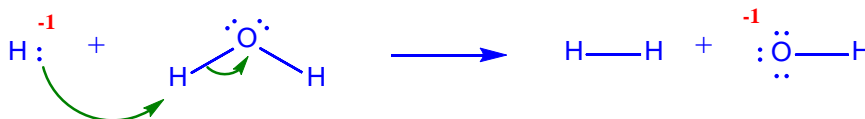
14. Complete the table below. Provide the best Lewis diagram for each compound, or one representative of the best set of Lewis diagrams for molecules with resonance structures.

Include any non-zero formal charges on the appropriate atom(s). **[8 marks]**

Note that hydrogen atoms are not usually attached to the central atom in oxoacids.

Formula	Name	Lewis diagram	Drawing of Molecule with Bond Angles Labeled
PO_3^{3-}	phosphite		
HNO_2	nitrous acid		

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15. Ethanol has a boiling point of 78 ˚C. Dimethyl ether has a boiling point of -24 ˚C. Given that these two compounds have the same molecular formula (C_2H_6O), explain why the boiling point of ethanol is so much higher than the boiling point of dimethyl ether.

[2 marks]

ethanol   dimethyl ether
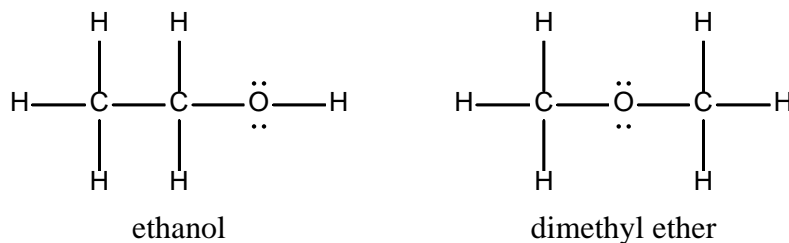
Both molecules are polar; however, ethanol molecules are able to hydrogen bond with other ethanol molecules (while dimethyl ether molecules cannot hydrogen bond with each other). As such, the intermolecular forces in liquid ethanol are stronger overall than those in liquid dimethyl ether and it has a higher boiling point.

Many students suggested that ethanol was polar and dimethyl ether was not. This is incorrect. Apply VSEPR theory to dimethyl ether and you will see that the oxygen atom has a bent geometry! That gives the molecule a net dipole, making it polar.

16.   **[6 marks]**

(a) Draw a Lewis diagram for a molecule with the molecular formula $C_2H_2F_2$ which is **polar**. *If the geometry of the molecule is not clear from your Lewis diagram, redraw the molecule to clearly show its geometry.*   *[2 marks]*

or

(b) Draw a Lewis diagram for a molecule with the molecular formula $C_2H_2F_2$ which is **nonpolar**. *If the geometry of the molecule is not clear from your Lewis diagram, redraw the molecule to clearly show its geometry.*
[2 marks]
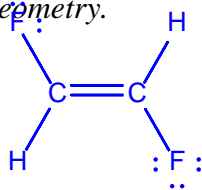
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17.   A 25 L container contains 2.25 kg of water vapour at 375 °C.   **[8 marks]**

(a)   What physical factors lead to deviations from the ideal gas law?   *[2 marks]*

An ideal gas has an extremely low density with the particles occupying a negligible fraction of the container volume and spread too far apart to experience any significant intermolecular forces.

As the density of a gas increases, the gas particles start to experience intermolecular forces pulling the gas particles toward each other. These intermolecular forces can be due to polarity (e.g. dipole-dipole attractions) as well as size (e.g. London dispersion forces).

As the density of a gas increases, the gas particles also start to occupy a larger fraction of the container volume. Since two gas particles cannot occupy the exact same position, this reduces the amount of 'empty space' in the container. This effect increases with the size of the gas particles.

(b)   Calculate the pressure in the container if the water vapour behaved as an ideal gas.
     [2 marks]

$$n_{H_2O} = 2.25kg \times \frac{1000g}{1kg} \times \frac{1mol}{18.0152g} = 125mol \qquad V = 25L \times \frac{1m^3}{1000L} = 0.025m^3$$

$$T = (375 + 273.15)K = 648K$$

$$PV = nRT$$

$$P = \frac{nRT}{V}$$

$$= \frac{(125mol)\left(8.314462\frac{Pa \cdot m^3}{mol \cdot K}\right)(648K)}{0.025m^3}$$

$$P = 2.7 \times 10^7 \, Pa = 27MPa$$

(c)   Calculate the actual pressure in the container given that the water vapour does not behave ideally.   *[4 marks]*

$$\left(P + a\frac{n^2}{V^2}\right)(V - nb) = nRT$$

$$P = \frac{nRT}{V - nb} - a\frac{n^2}{V^2}$$

$$= \frac{(125mol)\left(8.314462\frac{Pa \cdot m^3}{mol \cdot K}\right)(648K)}{(0.025m^3) - (125mol)\left(3.05 \times 10^{-5} \frac{m^3}{mol}\right)} - \left(0.5537 \frac{Pa \cdot m^6}{mol^2}\right)\frac{(125mol)^2}{(0.025m^3)^2}$$

$$= (3.2 \times 10^7 \, Pa) - (1.4 \times 10^7 \, Pa)$$

$$P = 1.8 \times 10^7 \, Pa = 18MPa$$

This is significantly lower than the pressure predicted by the ideal gas equation.

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18. In the 12 Amber Bottles lab, you made a dark red complex ion when you mixed potassium thiocyanate solution (KSCN) with iron(III) nitrate solution. The formula for this complex ion is $[Fe(SCN)(H_2O)_5]^x$ where x is the charge of the ion. **[11 marks]**

(a) What is the chemical formula for iron(III) nitrate? *[1 mark]*

$Fe(NO_3)_3$

(b) What is the overall charge of the complex ion? (In other words, what is x?) *[1 mark]*

+2

(c) What is the co-ordination number of iron in this complex ion? *[1 mark]*
Thiocyanate is a monodentate ligand.

6

(d) Write the electron configuration for iron(III). **Please use** a noble gas abbreviation. *[2 marks]*

[Ar] $3d^5$

(e) The complex ions of iron(III) in the reaction you performed are both in the high-spin state. Draw an energy-level diagram for the crystal field splitting of high-spin iron(III). Place the electrons and label the orbitals. *[3 marks]*
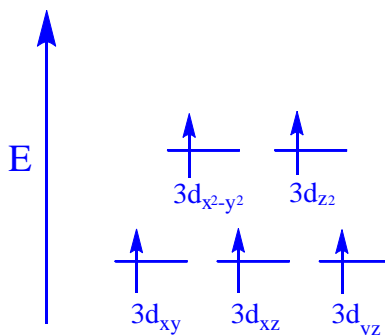
(f) **Briefly** explain why these iron complexes are coloured. *[3 marks]*

When light of the appropriate wavelength* is shone upon the complex, it is absorbed by an electron in one of the lower-energy 3d orbitals. This excites that electron into one of the higher-energy 3d orbitals. Since that colour of light is absorbed by the complex, it is removed from the spectrum of colours transmitted through the solution of complex (or reflected by a solid complex). The complex therefore appears to be the complement of the colour absorbed. e.g. If green light is absorbed, the complex will appear red.

*appropriate wavelength = light with energy exactly equal to the difference in energy between the lower-energy and higher-energy 3d orbitals

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19.   **[10 marks]**

(a) Why is hydrogen the only neutral atom for which the ionization energy can be readily calculated?   *[1 mark]*

Hydrogen is the only neutral atom with just one electron.

In atoms with more than one electron, there is more than just one force to factor into the calculation (electron-electron repulsions in addition to each electron-nucleus attraction).

(b) Calculate the ionization energy of a neutral ground-state hydrogen atom.   *[4 marks]*
Report your answer in J.

A hydrogen atom has 1 proton so $Z = 1$.

Ground state means that $n = 1$.

Ionization energy is the difference between the atom's initial energy and the energy if the electron left the atom ($n = \infty$; $E_{n=\infty} = 0\,J$).

$$E_i = E_{n=\infty} - E_{n=1} \qquad\qquad E_{n=1} = -R_H \frac{Z^2}{n^2}$$

$$E_{n=1} = -(2.179872 \times 10^{-18}\,J)\frac{(1)^2}{(1)^2}$$

$$E_{n=1} = -2.179872 \times 10^{-18}\,J$$

$$E_i = (0\,J) - (-2.179872 \times 10^{-18}\,J)$$
$$E_i = 2.179872 \times 10^{-18}\,J \qquad\qquad\qquad\qquad\qquad 7\text{ sig. fig.}$$

(c) Ionization energies are usually reported in kJ/mol. Convert your answer to part (b) into kJ/mol.   *[2 marks]*

$$E_i = 2.179872 \times 10^{-18}\,\frac{J}{atom} \times \frac{1\,kJ}{1000\,J} \times \frac{6.022141\times10^{23}\,atoms}{1\,mol} = 1312.750\,\frac{kJ}{mol} \qquad 7\text{ sig. fig.}$$

(d) Calculate the longest wavelength of light that could be used to ionize a hydrogen atom.
Report your answer in nm.   *[3 marks]*

One atom is ionized by one photon of light, so use the energy calculated in part (b).

$$E_{photon} = h\nu \qquad and \qquad\qquad c = \nu\lambda$$

$$E_{photon} = \frac{hc}{\lambda}$$

$$\lambda = \frac{hc}{E_{photon}} = \frac{\left(6.626070\times10^{-34}\,\frac{J}{Hz}\right)\left(2.997925\times10^{8}\,\frac{m}{s}\right)}{2.179872\times10^{-18}\,J} \times \frac{1\,Hz}{1\frac{1}{s}} \times \frac{10^9\,nm}{1\,m} = \mathbf{91.12673\,nm} \qquad 7\text{ sig.}$$
fig.

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20.   The following measurements were taken from a radioactive sample.   **[6 marks]**

Time (minutes)	Activity (Bq)
0	160
5	113
10	80
15	57
20	40

(a)   What is the half-life of this radioactive sample?   *[1 mark]*

10 minutes

This can be calculated using $ln\left(\frac{A_2}{A_1}\right) = -k(t_2 - t_1)$ *then* $ln(2) = k \cdot t_{1/2}$*; however it can also be observed directly from the data provided. The activity at 0 minutes is 160 Bq. After 10 minutes, the activity has halved to 80 Bq. After 20 minutes, the activity has halved again to 40 Bq. So, the half-life must be 10 minutes.*

(b)   Calculate the decay constant for this radioactive sample.   *[2 marks]*

$ln(2) = k \cdot t_{1/2}$ *therefore* $k = \frac{ln(2)}{t_{1/2}}$

$k = \frac{ln(2)}{t_{1/2}} = \frac{ln(2)}{10. \, min} = 0.069 \, min^{-1}$ 2 sig. fig.

(c)   At what time will the activity of the sample be 10 Bq?   *[1 mark]*

40 minutes

This can be calculated using $ln\left(\frac{A_2}{A_1}\right) = -k(t_2 - t_1)$*; however it can also be extrapolated directly from the data provided. The activity at 0 minutes is 160 Bq. After 10 minutes, the activity has halved to 80 Bq. After 20 minutes, the activity has halved again to 40 Bq. After 30 minutes, the activity should therefore half again to 20 Bq. After 40 minutes, the activity should half again to 10 Bq.*

(d)   How many radioactive nuclei were in the initial sample? (In other words, at time = 0)
          [2 marks]

$A = kN$ *therefore* $N = \frac{A}{k}$

$N = \frac{A}{k} = \frac{160 \, Bq}{0.069 \, min^{-1}} \times \frac{1 \, s^{-1}}{1 \, Bq} \times \frac{60 \, s}{1 \, min} = 1.4 \times 10^5$ 2 sig. fig.

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21. Small stars such as our sun convert hydrogen into helium via a series of reactions called the proton-proton chain reaction. **[6 marks]**

(a) In the first step of the proton-proton chain reaction, two ^{1}H nuclei react to give a positron and one other product. Write a balanced equation for this reaction. *[1 mark]*

$2{^1_1}H \rightarrow {^2_1}H + {^0_1}\beta$
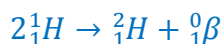
(b) What is a positron? *[1 mark]*

A positron is the antimatter equivalent of an electron.

(c) Calculate the energy released by this reaction. *[4 marks]*
Report your answer in kJ/mol.

$\Delta m = \sum m_{products} - \sum m_{products}$ *not including positrons or electrons*

$\Delta m = m_{hydrogen-2} - 2m_{hydrogen-1}$ *where 1 u = 1 g/mol*

$\Delta m = \left(2.014101778\frac{g}{mol}\right) - 2\left(1.007825032\frac{g}{mol}\right)$

$\Delta m = -0.001548286\frac{g}{mol}$ 9 decimal places so 7 sig. fig.

$\Delta E = \Delta mc^2$

$\Delta E = \left(-0.001548286\frac{g}{mol}\right)\left(2.997925 \times 10^8\frac{m}{s}\right)^2 \times \frac{1\,kg}{1000\,g} \times \frac{1\,J}{1\frac{kg\cdot m^2}{s^2}} \times \frac{1\,kJ}{1000\,J}$

$\Delta E = -1.391530 \times 10^8\frac{kJ}{mol}$ 7 sig. fig.

22. What was the most useful and/or interesting thing you learned in CHEM 1000? **[1 mark]**

**…AND THAT'S ALL FOR CHEM 1000.
HAPPY HOLIDAYS!**

DATA SHEET

Fundamental Constants and Conversion Factors

Atomic mass unit (u)	$1.660\,539 \times 10^{-27}$ kg	Kelvin temperature scale	0 K = -273.15 °C
Avogadro's number (N_A)	$6.022\,141 \times 10^{23}$ mol^{-1}	Planck's constant	$6.626\,070 \times 10^{-34}$ J·Hz^{-1}
Bohr radius (a_0)	$5.291\,772 \times 10^{-11}$ m	Proton mass	1.007 277 u
Electron charge (e)	$1.602\,177 \times 10^{-19}$ C	Neutron mass	1.008 665 u
Electron mass	$5.485\,799 \times 10^{-4}$ u	Rydberg Constant (R_H)	$2.179\,872 \times 10^{-18}$ J
Ideal gas constant (R)	$8.314\,462$ J·mol^{-1}·K^{-1}	Speed of light in vacuum	$2.997\,925 \times 10^8$ m·s^{-1}
	$8.314\,462$ m^3·Pa·mol^{-1}·K^{-1}	Standard atmospheric pressure	1 bar = 100 kPa
		Volume	1000 L = 1 m^3

Formulae

$$c = \lambda \upsilon \qquad E = h\upsilon \qquad p = mv \qquad \lambda = \frac{h}{p} \qquad \Delta x \cdot \Delta p > \frac{h}{4\pi} \qquad r_n = a_0 \frac{n^2}{Z} \qquad E_n = -R_H \frac{Z^2}{n^2}$$
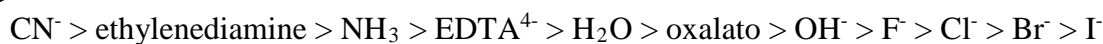
$$\overline{E_k} = \frac{1}{2}m\overline{v^2} = \frac{3}{2}\frac{RT}{N_A} \qquad v_{rms} = \sqrt{\overline{v^2}} = \sqrt{\frac{3RT}{M}} \qquad PV = nRT \qquad \left(P + a\frac{n^2}{V^2}\right)(V - bn) = nRT$$

$$\Delta E = \Delta mc^2 \qquad A = -\frac{\Delta N}{\Delta t} \qquad A = kN \qquad \ln\left(\frac{N_2}{N_1}\right) = -k(t_2 - t_1) \qquad \ln(2) = k \cdot t_{1/2}$$

$$pK_a \approx 8 - 5p \text{ for oxoacids } O_pE(OH)_q$$

Spectrochemical Series

strong field   weak field

$$CN^- > \text{ethylenediamine} > NH_3 > EDTA^{4-} > H_2O > \text{oxalato} > OH^- > F^- > Cl^- > Br^- > I^-$$

Band of Stability Graph
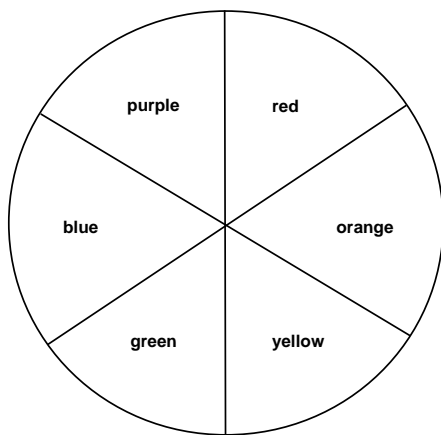
The graph at the right shows the band of stability. Stable isotopes are in black. Isotopes that exist but are not stable are shown in varying shades of gray with the shades of gray corresponding to different half-lives.

DATA SHEET

CHEM 1000 Standard Periodic Table

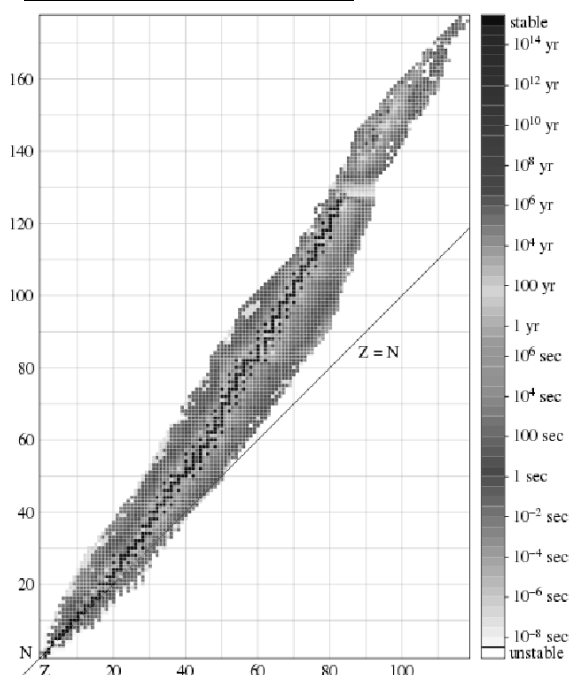| 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1.0079 **H** 1																	4.0026 **He** 2
6.941 **Li** 3	9.0122 4											10.811 **B** 5	12.011 **C** 6	14.0067 **N** 7	15.9994 **O** 8	18.9984 **F** 9	20.1797 **Ne** 10
22.9898 **Na** 11	24.3050 **Mg** 12											26.9815 **Al** 13	28.0855 **Si** 14	30.9738 **P** 15	32.066 **S** 16	35.4527 **Cl** 17	39.948 **Ar** 18
39.0983 **K** 19	40.078 **Ca** 20	44.9559 21	47.88 **Ti** 22	50.9415 **V** 23	51.9961 **Cr** 24	54.9380 **Mn** 25	55.847 **Fe** 26	58.9332 **Co** 27	58.693 28	63.546 **Cu** 29	65.39 **Zn** 30	69.723 **Ga** 31	72.61 **Ge** 32	74.9216 **As** 33	78.96 34	79.904 35	83.80 **Kr** 36
85.4678 **Rb** 37	87.62 **Sr** 38	88.9059 **Y** 39	91.224 **Zr** 40	92.9064 **Nb** 41	95.94 **Mo** 42	(98) **Tc** 43	101.07 **Ru** 44	102.906 **Rh** 45	106.42 **Pd** 46	107.868 **Ag** 47	112.411 **Cd** 48	114.82 **In** 49	118.710 **Sn** 50	121.757 **Sb** 51	127.60 **Te** 52	126.905 **I** 53	131.29 **Xe** 54
132.905 **Cs** 55	137.327 **Ba** 56	**La-Lu**	178.49 **Hf** 72	180.948 **Ta** 73	183.85 **W** 74	186.207 **Re** 75	190.2 **Os** 76	192.22 **Ir** 77	195.08 **Pt** 78	196.967 **Au** 79	200.59 **Hg** 80	204.383 **Tl** 81	207.19 **Pb** 82	208.980 **Bi** 83	(210) **Po** 84	(210) **At** 85	(222) **Rn** 86
(223) **Fr** 87	226.025 **Ra** 88	**Ac-Lr**	(265) **Rf** 104	(268) **Db** 105	(271) **Sg** 106	(270) **Bh** 107	(277) **Hs** 108	(276) **Mt** 109	(281) **Ds** 110	(280) **Rg** 111	(285) **Cn** 112	(284) **Uut** 113	(289) **Fl** 114	(288) **Uup** 115	(293) **Lv** 116	(294) **Uus** 117	(294) **Uuo** 118

138.906 **La** 57	140.115 **Ce** 58	140.908 **Pr** 59	144.24 **Nd** 60	(145) **Pm** 61	150.36 **Sm** 62	151.965 **Eu** 63	157.25 **Gd** 64	158.925 **Tb** 65	162.50 **Dy** 66	164.930 **Ho** 67	167.26 **Er** 68	168.934 **Tm** 69	173.04 **Yb** 70	174.967 **Lu** 71
227.028 **Ac** 89	232.038 **Th** 90	231.036 **Pa** 91	238.029 **U** 92	237.048 **Np** 93	(240) **Pu** 94	(243) **Am** 95	(247) **Cm** 96	(247) **Bk** 97	(251) **Cf** 98	(252) **Es** 99	(257) **Fm** 100	(258) **Md** 101	(259) **No** 102	(262) **Lr** 103

Developed by Prof. R. T. Boeré (updated 2014)

Some Useful Masses

^1_1H	1.007 825 032 u
^2_1H	2.014 101 778 u
^3_2He	3.016 029 319 u
^4_2He	4.002 603 254 u
$^4_2\alpha$	4.001 506 179 u
^1_1p	1.007 276 467 u
^1_0n	1.008 664 916 u

van der Waals constants for $H_2O_{(g)}$

$a = 0.5537\ \text{Pa·m}^6\text{·mol}^{-2}$

$b = 3.05 \times 10^{-5}\ \text{m}^3\text{·mol}^{-1}$