

**AP<sup>®</sup> CHEMISTRY**  
**2007 SCORING GUIDELINES (Form B)**

**Question 2**

Answer the following problems about gases.

- (a) The average atomic mass of naturally occurring neon is 20.18 amu. There are two common isotopes of naturally occurring neon as indicated in the table below.

Isotope	Mass (amu)
Ne-20	19.99
Ne-22	21.99

- (i) Using the information above, calculate the percent abundance of each isotope.

<p>Let <math>x</math> represent the natural abundance of Ne-20.</p> $19.99x + 21.99(1-x) = 20.18$ $19.99x + 21.99 - 21.99x = 20.18$ $19.99x - 21.99x = 20.18 - 21.99$ $-2x = -1.81$ $x = 0.905$ <p><math>\Rightarrow</math> percent abundances are:    Ne-20 = 90.5%            Ne-22 = 9.5%</p>	<p>One point is earned for the correct answer.</p>
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- (ii) Calculate the number of Ne-22 atoms in a 12.55 g sample of naturally occurring neon.

$12.55 \text{ g Ne} \times \frac{1 \text{ mol Ne}}{20.18 \text{ g Ne}} \times \frac{0.095 \text{ mol Ne-22}}{1 \text{ mol Ne}} \times \frac{6.022 \times 10^{23} \text{ Ne-22 atoms}}{1 \text{ mol Ne-22}}$ $= 3.6 \times 10^{22} \text{ Ne-22 atoms}$	<p>One point is earned for the correct molar mass.</p> <p>One point is earned for the correct fraction of Ne-22 in Ne.</p> <p>One point is earned for the number of atoms.</p>
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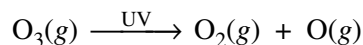
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**Question 2 (continued)**

- (b) A major line in the emission spectrum of neon corresponds to a frequency of  $4.34 \times 10^{14} \text{ s}^{-1}$ . Calculate the wavelength, in nanometers, of light that corresponds to this line.

$c = \lambda \nu \Rightarrow \lambda = \frac{c}{\nu}$ $\lambda = \frac{3.0 \times 10^8 \text{ m s}^{-1}}{4.34 \times 10^{14} \text{ s}^{-1}} \times \frac{1 \text{ nm}}{10^{-9} \text{ m}} = 690 \text{ nm}$	<p>One point is earned for the correct setup.</p> <p>One point is earned for the answer.</p>
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- (c) In the upper atmosphere, ozone molecules decompose as they absorb ultraviolet (UV) radiation, as shown by the equation below. Ozone serves to block harmful ultraviolet radiation that comes from the Sun.



A molecule of  $\text{O}_3(\text{g})$  absorbs a photon with a frequency of  $1.00 \times 10^{15} \text{ s}^{-1}$ .

- (i) How much energy, in joules, does the  $\text{O}_3(\text{g})$  molecule absorb per photon?

$E = h\nu$ $= 6.63 \times 10^{-34} \text{ J s} \times 1.00 \times 10^{15} \text{ s}^{-1}$ $= 6.63 \times 10^{-19} \text{ J per photon}$	<p>One point is earned for the correct answer.</p>
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- (ii) The minimum energy needed to break an oxygen-oxygen bond in ozone is  $387 \text{ kJ mol}^{-1}$ . Does a photon with a frequency of  $1.00 \times 10^{15} \text{ s}^{-1}$  have enough energy to break this bond? Support your answer with a calculation.

$\frac{6.63 \times 10^{-19} \text{ J}}{1 \text{ photon}} \times \frac{6.022 \times 10^{23} \text{ photons}}{1 \text{ mol}} \times \frac{1 \text{ kJ}}{10^3 \text{ J}} = 399 \text{ kJ mol}^{-1}$ <p><math>399 \text{ kJ mol}^{-1} &gt; 387 \text{ kJ mol}^{-1}</math>, therefore the bond can be broken.</p>	<p>One point is earned for calculating the energy.</p> <p>One point is earned for the comparison of bond energies.</p>
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