

Name: \_\_\_\_\_ Student Number: \_\_\_\_\_

## CHEM 1001 A and T Midterm Test #1

October 13, 2017

Calculators Allowed

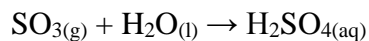
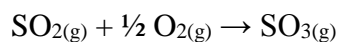
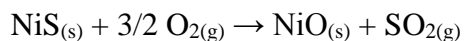
Make sure this test has 7 pages. You may tear off the last page.

**Part A. Answer each of the six questions with a few sentences or equations where necessary.  
(5 Marks each)****A1.** Name and explain two reasons why gases do not obey the ideal gas law at high pressures.

One reason is that molecules occupy volume themselves. This tends to increase the pressure above what the ideal gas law predicts. Another reason is that molecules attract one another, effectively reducing the pressure.

**A2.** Why does the temperature of the atmosphere go through a maximum at an altitude of 30 km?

The concentration of ozone is at a maximum at this altitude. Ozone absorbs infrared radiation and causes warming.

**A3.** Beginning with NiS<sub>(s)</sub>, show three reactions that lead to acid rain.

**A4.** Why does a constant pressure calorimeter not give a direct measurement of the internal energy change of a reaction?

A measurement at constant pressure yields the enthalpy change of the system. This is different from the internal energy change of the system by an amount  $p\Delta V$ .

**A5.** Why can an electron move from one lobe of an orbital to another without actually existing at the node between these two lobes?

The electron can also behave as a wave. A wave can have zero amplitude at the node, i.e. the electron does not exist at the node.

**A6.** Name three possible methods for storing hydrogen for use as a fuel in automobiles.

Any three of:

- Stored as a compressed gas
- Stored as a liquid
- Stored in carbon nanotubes
- Stored in a metal such as Palladium

**Part B. Answer any three of the following four questions (B1, B2, B3, B4). If you answer all four, the best three answers will count. (20 marks each)**

**B1.** (a) Chlorine gas,  $\text{Cl}_{2(\text{g})}$ , is produced from seawater via the “chlor-alkali” process. The gas is stored in containers to prevent unwanted and explosive reactions. If a 10.0 L container holds 8.17 mol  $\text{Cl}_{2(\text{g})}$  at  $200^\circ\text{C}$ :

(a) Calculate the pressure (in bar) assuming ideal behaviour.

$$p = \frac{nRT}{V} = \frac{8.17 \text{ mol} \times 0.08314 \text{ L bar K}^{-1} \text{ mol}^{-1} \times (200 + 273) \text{ K}}{10.0 \text{ L}}$$

$$= 32.1 \text{ bar}$$

(b) Calculate the pressure (in bar) using the van der Waals equation. For  $\text{Cl}_{2(\text{g})}$ ,  $a = 6.58 \text{ bar L}^2 \text{ mol}^{-2}$  and  $b = 0.0562 \text{ L mol}^{-1}$ .

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

$$= \frac{8.17 \text{ mol} \times 0.08314 \text{ L bar K}^{-1} \text{ mol}^{-1} \times (200 + 273) \text{ K}}{10.0 \text{ L} - 8.17 \text{ mol} \times 0.0562 \text{ L mol}^{-1}} - 6.58 \text{ bar L}^2 \text{ mol}^{-2} \left( \frac{8.17 \text{ mol}}{10.0 \text{ L}} \right)^2$$

$$= 33.67 \text{ bar} - 4.39 \text{ bar}$$

$$= 29.3 \text{ bar}$$

(c) Why is the van der Waals pressure lower than the ideal gas pressure for chlorine under these conditions?

The van der Waals equation takes account of the fact that the chlorine molecules have volume and interact. In this case, the intermolecular forces are reducing the pressure below what the ideal gas law predicts.

(d) Repeat the calculation in part (b) using the same number of moles of hydrogen in the same volume at the same temperature. For hydrogen,  $a = 0.245 \text{ bar L}^2 \text{ mol}^{-2}$  and  $b = 0.0265 \text{ L mol}^{-1}$ .

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

$$= \frac{8.17 \text{ mol} \times 0.08314 \text{ L bar K}^{-1} \text{ mol}^{-1} \times (200 + 273) \text{ K}}{10.0 \text{ L} - 8.17 \text{ mol} \times 0.0265 \text{ L mol}^{-1}} - 0.245 \text{ bar L}^2 \text{ mol}^{-2} \left( \frac{8.17 \text{ mol}}{10.0 \text{ L}} \right)^2$$

$$= 32.84 \text{ bar} - 0.16 \text{ bar}$$

$$= 32.7 \text{ bar}$$

(e) Why is the calculated pressure for hydrogen higher than that for chlorine?

Both the  $a$  and  $b$  terms are smaller for  $\text{H}_2$  than for  $\text{Cl}_2$ . In this case the pressure of  $\text{H}_2$  is higher than that of  $\text{Cl}_2$  so apparently the effect of its ‘ $b$ ’ term is greater than that of its ‘ $a$ ’ term.

- B2.** (a) The density of an unknown gas is found to be 1.29 g/L at 10.0 bar and 100°C. Calculate the molecular weight of the gas (g/mol), and identify the gas.

$$\rho = \frac{pM}{RT} \quad \text{thus, } M = \frac{\rho RT}{p} = \frac{1.29 \text{ g L}^{-1} (0.08314 \frac{\text{L bar}}{\text{mol K}}) (100 + 273) \text{ K}}{10.0 \text{ bar}} = 4.00 \text{ g mol}^{-1}$$

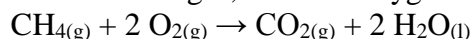
Only He has this molar mass.

- (b) Calculate the ratio of the diffusion rates of  $\text{SO}_{2(\text{g})}$  and  $\text{SO}_{3(\text{g})}$ . Which gas diffuses faster?

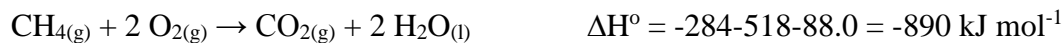
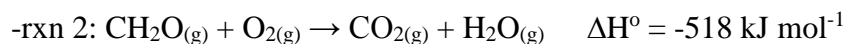
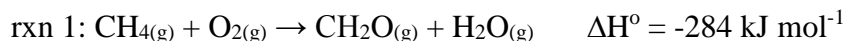
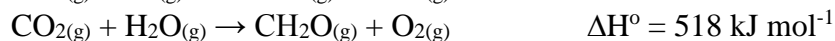
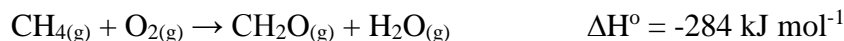
$$\frac{\text{rate}_{\text{SO}_2}}{\text{rate}_{\text{SO}_3}} = \sqrt{\frac{M_{\text{SO}_3}}{M_{\text{SO}_2}}} = \sqrt{\frac{80.1 \text{ g mol}^{-1}}{64.1 \text{ g mol}^{-1}}} = 1.11 \quad (\text{would be 0.90 if inverted, which is also fine})$$

The  $\text{SO}_2$  diffuses faster.

- (c) Methane, the main constituent of natural gas, burns in oxygen in the reaction:



Calculate  $\Delta H^\circ$  (in  $\text{kJ mol}^{-1}$ ) for this reaction using the following information:



**B3.** Some data are given for the reaction  $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

	$\text{CaCO}_3(\text{s})$	$\text{CaO}(\text{s})$	$\text{CO}_2(\text{g})$
$\Delta H_f^\circ$ , kJ mol <sup>-1</sup>	-1,207.6	-634.9	-393.5

Assuming 1000 kg of  $\text{CaCO}_3(\text{s})$  reacts at 1000°C and 1.00 bar:

(a) Calculate the total enthalpy change (in kJ)

$$\begin{aligned}\Delta H^\circ &= \Delta H_f^\circ(\text{CaO}(\text{s})) + \Delta H_f^\circ(\text{CO}_2(\text{g})) - \Delta H_f^\circ(\text{CaCO}_3(\text{s})) \\ &= -634.9 + (-393.5) - (-1207.6) \\ &= 179.2 \text{ kJ mol}^{-1}\end{aligned}$$

$$1000 \text{ kg CaCO}_3 \times \left(\frac{1000 \text{ g}}{1 \text{ g}}\right) \times \left(\frac{1 \text{ mol}}{100.1 \text{ g}}\right) \times \left(\frac{179.2 \text{ kJ}}{1 \text{ mol}}\right) = 1.79 \times 10^6 \text{ kJ}$$

(b) Calculate the total work done (in kJ)

$$1000 \text{ kg CaCO}_3 \times \left(\frac{1000 \text{ g}}{1 \text{ g}}\right) \times \left(\frac{1 \text{ mol CaCO}_3}{100.1 \text{ g CaCO}_3}\right) \times \left(\frac{1 \text{ mol CO}_2}{1 \text{ mol CaCO}_3}\right) = 9.99 \times 10^3 \text{ mol CO}_2(\text{g})$$

$$\begin{aligned}V &= \frac{nRT}{p} \\ &= \frac{9.99 \times 10^3 \text{ mol}(0.08314 \text{ L bar K}^{-1} \text{ mol}^{-1})(1000 + 273) \text{ K}}{1.00 \text{ bar}}\end{aligned}$$

$$= 1.06 \times 10^6 \text{ L}$$

$$= 1.06 \times 10^3 \text{ m}^3$$

$$w = -p\Delta V$$

$$= -10^5 \text{ Pa}(1.06 \times 10^3 \text{ m}^3)$$

$$= -1.06 \times 10^8 \text{ Pa m}^3$$

$$= -1.06 \times 10^8 \text{ J}$$

$$= -1.06 \times 10^5 \text{ kJ}$$

(c) Calculate the total  $\Delta E$  (in kJ). Has the system gained or lost energy?

$$\begin{aligned}\Delta E &= q + w \\ &= 1.79 \times 10^6 \text{ kJ} - 1.06 \times 10^5 \text{ kJ} \\ &= 1.68 \times 10^6 \text{ kJ}\end{aligned}$$

The system has gained this energy ( $\Delta E > 0$ ).

- B4.** (a) The Balmer-Rydberg equation can also be used for  $\text{He}^+$  because it has one electron (the same as an H atom). The lines in the spectrum of  $\text{He}^+$  from stars include transitions to the  $n = 4$  level. Calculate the energy (in  $\text{kJ mol}^{-1}$ ) of the lowest energy transition in this series.

$$\begin{aligned}\frac{1}{\lambda} &= R \left[ \frac{1}{m^2} - \frac{1}{n^2} \right] \\ &= 0.01097 \text{ nm}^{-1} \left[ \frac{1}{4^2} - \frac{1}{5^2} \right] \\ &= 2.47 \times 10^{-4} \text{ nm}^{-1} \\ \lambda &= \frac{1}{2.47 \times 10^{-4} \text{ nm}^{-1}} = 4051 \text{ nm}\end{aligned}$$

$$\begin{aligned}E &= h\nu \\ &= h \left( \frac{c}{\lambda} \right) \\ &= 6.63 \times 10^{-34} \text{ J s} \left( \frac{3.00 \times 10^8 \text{ m s}^{-1}}{4051 \times 10^{-9} \text{ m}} \right) \\ &= 4.91 \times 10^{-20} \text{ J}\end{aligned}$$

$$\begin{aligned}\times 6.02 \times 10^{23} \text{ mol}^{-1} &= 2.96 \times 10^4 \text{ J mol}^{-1} \\ &= 29.6 \text{ kJ mol}^{-1}\end{aligned}$$

- (b) The binding energy of electrons to a chromium metal surface is  $7.20 \times 10^{-19} \text{ J}$ . Calculate the wavelength (in nm) of light required to give ejected electrons a kinetic energy of  $2.50 \times 10^{-19} \text{ J}$ .

The incoming photon must supply the binding energy plus the energy of the ejected electron. In this case this sum is  $7.20 \times 10^{-19} + 2.50 \times 10^{-19} = 9.70 \times 10^{-19} \text{ J}$ .

$$\begin{aligned}E &= h\nu = h \left( \frac{c}{\lambda} \right) \\ \lambda &= \frac{hc}{E} \\ &= \frac{6.63 \times 10^{-34} \text{ J s} (3.00 \times 10^8 \text{ m s}^{-1})}{9.70 \times 10^{-19} \text{ J}} \\ &= 2.05 \times 10^{-7} \text{ m} \\ &= 205 \text{ nm}\end{aligned}$$

A	B1	B2	B3	B4	Total/90