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## Final Exam – CHM 1311-F

**Date:** \_\_\_\_\_ **Length:** \_\_\_\_\_

**Last Name:**

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**First**

**Student #** \_\_\_\_\_ **Seat #**

### Instructions:

- Calculators of any kind are permitted without WiFi capability.
- Closed book exam.
- 3 hrs exam.
- Periodic table and table of electronegativity are allowed.
- Relevant data and equations are at the end.
- This booklet contains # pages.

Cellular phones, unauthorized electronic devices or course notes (unless an open-book exam) are not allowed during this exam. Phones and devices must be turned off and put away in your bag. Do not keep them in your possession, such as in your pockets. If caught with such a device or document, the following may occur: you will be asked to leave immediately the exam, academic fraud allegations will be filed which may result in you obtaining a 0 (zero) for the exam.

### Read carefully:

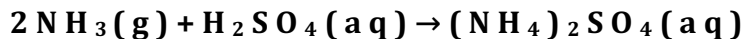
By signing below, you acknowledge that you have read and ensured that you are complying with the above statement.

Signature: \_\_\_\_\_

Please put your INITIALS IN THE BOX when you have verified that there are # pages in this exam.

1. (1 point)

The fertilizer ammonium sulphate is prepared by the reaction between ammonia and sulphuric acid:



What mass (in kg) of  $\text{NH}_3$  is needed to make 3.50 tonnes (1 tonne = 1000 kg) of ammonium sulphate?

You are asked to calculate the mass of ammonia required to produce 3.50 metric tons of ammonium sulphate using the chemical equation:  $2 \text{NH}_3 + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4$ . The result is asked for in kilograms; thus, the kg to g mass unit conversions will cancel and do not need to be shown:

$$3.50 \text{ ton } (\text{NH}_4)_2\text{SO}_4 \left( \frac{10^3 \text{ kg}}{1 \text{ ton}} \right) \left( \frac{1 \text{ mol}}{132.1 \text{ g}} \right) \left( \frac{2 \text{ mol NH}_3}{1 \text{ mol } (\text{NH}_4)_2\text{SO}_4} \right) \left( \frac{17.04 \text{ g}}{1 \text{ mol}} \right) = 903 \text{ kg NH}_3$$

2. (1 point)

The chemical formula for aluminum sulphate is  $\text{Al}_2(\text{SO}_4)_3$ . (a) Compute its molar mass. (b) Compute the number of moles contained in 25.0 g of this compound. (c) Determine its percent composition. (d) Determine the mass of this compound that contains 1.00 mole of O.

The chemical formula of a substance provides all the information needed to compute its molar characteristics:

(a)  $M = 2(26.98 \text{ g/mol}) + 3[32.07 \text{ g/mol} + 4(16.00 \text{ g/mol})] = 342.17 \text{ g/mol}$

(b)  $n = \frac{m}{M} = 25.0 \text{ g} \left( \frac{1 \text{ mol}}{342.17 \text{ g}} \right) = 7.31 \times 10^{-2} \text{ mol}$

(c) To determine the percent composition, work with 1 mole of substance. Take the ratio of the mass of each element that 1 mole contains to the mass of 1 mole (molar mass):

$$\%Al = (100\%) \left( \frac{2(26.98 \text{ g/mol})}{342.17 \text{ g/mol}} \right) = 15.77\%$$

$$\%S = (100\%) \left( \frac{3(32.07 \text{ g/mol})}{342.17 \text{ g/mol}} \right) = 28.12\%$$

$$\%O = (100\%) \left( \frac{12(16.00 \text{ g/mol})}{342.17 \text{ g/mol}} \right) = 56.11\%$$

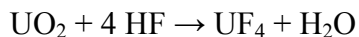
$$(d) 1.00 \text{ mol O} \left( \frac{1 \text{ mol Al}_2(\text{SO}_4)_3}{12 \text{ mol O}} \right) \left( \frac{342.17 \text{ g}}{1 \text{ mol}} \right) = 28.5 \text{ g}$$

**3. ( 2 points)**

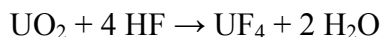
**Uranium tetrafluoride (UF<sub>4</sub>) is an intermediate in the production of uranium hexafluoride (UF<sub>6</sub>). One method of manufacturing UF<sub>4</sub> consists of reacting UO<sub>2</sub> with HF according to the (unbalanced) reaction UO<sub>2</sub> + HF → UF<sub>4</sub> + H<sub>2</sub>O. Assuming a yield of 99%, calculate the volume of 31.8 M HF required to manufacture 25.0 tonnes of UF<sub>4</sub>**

First we need a balanced reaction to work with. Because there are four F atoms on the right, we need four

HF molecules on the left:



Now the H and O atoms can be balanced by increasing the coefficient of H<sub>2</sub>O to 2:



25.0 tonnes of UF<sub>4</sub> per day is converted into moles:

$$\frac{25.0 \text{ tonnes UF}_4}{\text{day}} \times \frac{1000 \text{ kg}}{1 \text{ tonne}} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol}}{314 \text{ g}} = \frac{79600 \text{ mol UF}_4}{\text{day}}$$

**Notice:** Just ignore the “day”

From the balanced reaction, we see that the manufacture of 1 mol UF<sub>4</sub> requires 4 mol HF. Accounting for the yield, the total HF required is therefore

$$\frac{79600 \text{ mol UF}_4}{\text{day}} \times \frac{4 \text{ mol HF}}{1 \text{ mol UF}_4} \times \frac{100\%}{99\%} = \frac{322,000 \text{ mol HF}}{\text{day}}$$

Now we can calculate the required volume of HF solution:

$$\frac{322,000 \text{ mol HF}}{\text{day}} \times \frac{1 \text{ L}}{31.8 \text{ mol HF}} = \frac{10,100 \text{ L}}{\text{day}}$$

**4. (2 points)**

**Sodium metal reacts with molecular chlorine gas to form sodium chloride. A closed container of volume  $3.00 \times 10^3$  mL contains chlorine gas at 27 °C and  $1.25 \times 10^3$  Torr. Then, 6.90 g of solid sodium is introduced, and the reaction goes to completion. What is the final pressure (in bar) if the temperature rises to 47 °C?**

This is a stoichiometry problem that involves gases. We are asked to determine the final pressure of the container (which is the pressure of chlorine gas). Begin by analyzing the chemistry. The starting materials are Na metal and Cl<sub>2</sub>, a gas. The product is NaCl. The balanced chemical reaction is



The problem gives information about the amounts of both starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the final pressure. Calculations of initial amounts:

$$n_{\text{Na}} = 6.90 \text{ g} \left( \frac{1 \text{ mol}}{22.99 \text{ g}} \right) = 0.300 \text{ mol}$$

For chlorine gas, use the ideal gas equation to do pressure–mole conversions. The data must have the same units as those for *R*:

$$p_{\text{Cl}_2} = 1.25 \times 10^3 \text{ Torr} \left( \frac{1 \text{ bar}}{750.06 \text{ Torr}} \right) = 1.67 \text{ bar}$$

$$n_{\text{Cl}_2} = \frac{pV}{RT} = \frac{(1.67 \text{ bar})(3.00 \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(300 \text{ K})} = 0.200 \text{ mol}$$

Divide each initial amount by its coefficient to determine the limiting reactant:

$$\text{Cl}_2: 0.200 \text{ mol} \quad \text{Na: } \frac{0.300 \text{ mol}}{2} = 0.150 \text{ mol (LR)}$$

Use the initial amounts and the balanced equation to construct a table of amounts:

Reaction:	2 Na (s)	+ Cl <sub>2</sub> (g)	→ 2 NaCl (s)
Initial amount (mol)	0.300	0.200	0.000
Change (mol)	-0.300	-0.150	+0.300
Final amount (mol)	0.000	0.050	0.300

Use the final amount of Cl<sub>2</sub> and the new temperature (47 °C) to calculate the pressure:

$$p = \frac{nRT}{V} = \frac{(5.0 \times 10^{-2} \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(320 \text{ K})}{3.00 \text{ L}} = 0.443 \text{ bar}$$

### 5. (3 points)

**Freons (CFCs) are compounds that contain carbon, chlorine, and fluorine in various proportions. They have been extensively used as foaming agents, propellants, and refrigeration fluids. Freons are controversial because of the damage they do to the ozone layer in the stratosphere. A 2.55 g sample of a particular Freon in a 1.50 L bulb at 25.0 °C has a pressure of 262 Torr. What are the molar mass and formula of the compound?**

The ideal gas equation,  $pV = nRT$ , can be used to calculate moles using  $p$ - $V$ - $T$  data. Molar mass is related

to moles through  $n = \frac{m}{M}$ :

$$n = \frac{pV}{RT} = \frac{m}{M} \quad M = \frac{mRT}{pV}$$

Begin by converting the initial data into the units of  $R$ :

$$T = 25.0 + 273.15 = 298. \text{ K}$$

$$p = 262 \text{ Torr} \left( \frac{1 \text{ bar}}{750.06 \text{ Torr}} \right) = 0.349 \text{ bar}$$

Use the modified ideal gas equation to obtain the molar mass:

$$M = \frac{(2.55 \text{ g})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(298 \text{ K})}{(0.349 \text{ bar})(1.50 \text{ L})} = 121 \text{ g/mol}$$

The compound contains only C (12.0 g/mol), F (19.0 g/mol), and Cl (35.5 g/mol). The formula can be determined by trial and error. The combination of 1 C, 2 F, and 2 Cl has

$$M = 1(12.0 \text{ g/mol}) + 2(19.0 \text{ g/mol}) + 2(35.5 \text{ g/mol}) = 121 \text{ g/mol}$$

This matches the experimental value. The formula is  $\text{CF}_2\text{Cl}_2$ .

### 6. (1 point)

**A mouse is placed in a sealed chamber filled with air at 765 Torr and equipped with enough solid KOH to absorb any  $\text{CO}_2$  and  $\text{H}_2\text{O}$  produced. The gas volume in the chamber is 2.05 L, and its temperature is held at 298 K. After 2 hours, the pressure inside the chamber has fallen to 725 Torr. What mass of oxygen has the mouse consumed?**

As the mouse breathes, it inhales air containing oxygen. Some of this oxygen is used for metabolism, resulting in  $\text{CO}_2$ , which the mouse exhales. The amount of oxygen consumed is therefore equal to the amount of  $\text{CO}_2$  exhaled. The solid KOH in the chamber absorbs all the  $\text{CO}_2$  and any water the mouse exhales. Any reduction in pressure is therefore due to the removal of oxygen from the atmosphere.

The initial amount of gas in the chamber is

$$n_{\text{initial}} = \frac{p_{\text{initial}}V}{RT} = \frac{(765 \text{ Torr}) \left( \frac{1 \text{ bar}}{750.06 \text{ Torr}} \right) (2.05 \text{ L})}{(0.08314 \text{ L bar K}^{-1} \text{ mol}^{-1})(298 \text{ K})} = 0.08439 \text{ mol}$$

And after two hours, the amount of gas remaining is

$$n_{\text{initial}} = \frac{P_{\text{initial}}V}{RT} = \frac{(725 \text{ Torr})\left(\frac{1 \text{ bar}}{750.06 \text{ Torr}}\right)(2.05 \text{ L})}{(0.08314 \text{ L bar K}^{-1}\text{mol}^{-1})(298 \text{ K})} = 0.07998 \text{ mol}$$

The amount of oxygen consumed is therefore  $0.08439 - 0.07998 \text{ mol} = 0.00441 \text{ mol}$ .

$$(0.00441 \text{ mol})(32.00 \text{ g/mol}) = 0.141 \text{ g O}_2.$$

### 7. (3 points)

**An iron kettle weighing 1.35 kg contains 2.75 kg of water at 23.0 °C. The kettle and water are heated to 95.0 °C. How many joules of energy are absorbed by the water and by the kettle? (molar  $c_{\text{H}_2\text{O}} = 75.291 \text{ J mol}^{-1} \text{ }^\circ\text{C}^{-1}$ ; molar  $c_{\text{Fe}} = 25.10 \text{ J mol}^{-1} \text{ }^\circ\text{C}^{-1}$ )**

$$q = nC\Delta T \qquad \Delta T = 95.0 - 23.0 = 72.0 \text{ }^\circ\text{C}$$

$$n_{\text{kettle}} = 1.35 \text{ kg} \left(\frac{1000 \text{ g}}{1 \text{ kg}}\right) \left(\frac{1 \text{ mol}}{55.85 \text{ g}}\right) = 24.2 \text{ mol Fe}$$

$$q_{\text{kettle}} = (24.2 \text{ mol})(25.10 \text{ J mol}^{-1} \text{ }^\circ\text{C}^{-1})(72.0 \text{ }^\circ\text{C}) = 4.37 \times 10^4 \text{ J}$$

$$n_{\text{water}} = 2.75 \text{ kg} \left(\frac{1000 \text{ g}}{1 \text{ kg}}\right) \left(\frac{1 \text{ mol}}{18.02 \text{ g}}\right) = 153 \text{ mol water}$$

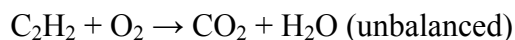
$$q_{\text{water}} = (153 \text{ mol})(75.291 \text{ J mol}^{-1} \text{ }^\circ\text{C}^{-1})(72.0 \text{ }^\circ\text{C}) = 8.29 \times 10^5 \text{ J}$$

**Alternatively the molar heat capacities can be converted to per-gram heat capacities and grams instead of moles can be used for the same calculation.**

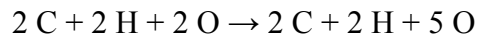
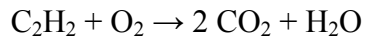
### 8. (2 points)

**Acetylene ( $\text{C}_2\text{H}_2$ ) is used in welding torches because it has a high enthalpy of combustion. When 1.00 g of acetylene burns completely in excess  $\text{O}_2$  gas at constant volume, it releases 48.2 kJ of energy. (a) What is the balanced chemical equation for this reaction? (b) What is the molar energy of combustion of acetylene? (c) How much energy is released per mole of  $\text{O}_2$  consumed?**

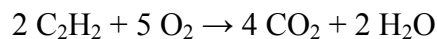
In a combustion reaction, the products are  $\text{CO}_2$  and  $\text{H}_2\text{O}$ :



Follow standard procedures to balance the equation. Give  $\text{CO}_2$  a coefficient of 2 to balance C:



Give  $\text{O}_2$  a coefficient of  $5/2$  to balance O, and then multiply by 2 to clear fractions:



(b) Find energy per mole from the energy released by 1.00 g using the molar mass (26.04 g/mol):

$$\Delta E = \left( \frac{-48.2 \text{ kJ}}{1.00 \text{ g}} \right) \left( \frac{26.04 \text{ g}}{1 \text{ mol}} \right) = -1.26 \times 10^3 \text{ kJ/mol C}_2\text{H}_2$$

(c) Five moles of  $\text{O}_2$  is consumed for every 2 moles of acetylene, so the energy released per mole of  $\text{O}_2$  is

$$\Delta E = \left( \frac{-1.26 \times 10^3 \text{ kJ}}{1 \text{ mol acetylene}} \right) \left( \frac{2 \text{ mol acetylene}}{5 \text{ mol O}_2} \right) = -5.04 \times 10^2 \text{ kJ/mol O}_2$$

### 9. (4 points)

**When light of frequency  $1.30 \times 10^{15} \text{ s}^{-1}$  shines on the surface of cesium metal, electrons are ejected with a maximum kinetic energy of  $5.2 \times 10^{-19} \text{ J}$ . Calculate (a) the wavelength of this light; (b) the binding energy of electrons to cesium metal; and (c) the longest wavelength of light that will eject electrons.**

$$4.17 \text{ (a)} \quad \lambda = \frac{c}{\nu} = \frac{2.998 \times 10^8 \text{ m/s}}{1.30 \times 10^{15} \text{ s}^{-1}} = 2.31 \times 10^{-7} \text{ m}$$

(b) First determine the energy of the photon, then use  $E_{\text{binding}} = E_{\text{photon}} - E_{\text{kinetic}}$

$$E_{\text{photon}} = h\nu = (6.626 \times 10^{-34} \text{ J s})(1.30 \times 10^{15} \text{ s}^{-1}) = 8.61 \times 10^{-19} \text{ J}$$

$$E_{\text{binding}} = (8.61 \times 10^{-19} \text{ J}) - (5.2 \times 10^{-19} \text{ J}) = 3.4 \times 10^{-19} \text{ J}$$

(c) The longest wavelength that will eject electrons corresponds to a photon with energy equal to the

$$\text{binding energy: } \lambda = \frac{hc}{E_{\text{binding}}}$$

$$\lambda = \frac{(6.626 \times 10^{-34} \text{ J s})(2.998 \times 10^8 \text{ m/s})}{3.4 \times 10^{-19} \text{ J}} = 5.8 \times 10^{-7} \text{ m}$$

**10. (1 point)**

**Determine the frequencies that hydrogen atoms emit in transitions from the  $n = 6$  and  $n = 5$  levels to the  $n = 3$  level. In what region of the electromagnetic spectrum do these photons lie? ( $E_n = -2.18 \times 10^{-18} / n^2$ )**

Energies and frequencies for transitions in hydrogen atoms can be calculated from the equation for hydrogen atom energy levels:

$$E_n = \frac{-2.18 \times 10^{-18} \text{ J}}{n^2}$$

$$\Delta E_{6-3} = E_6 - E_3 = \left( \frac{-2.18 \times 10^{-18} \text{ J}}{6^2} \right) - \left( \frac{-2.18 \times 10^{-18} \text{ J}}{3^2} \right) = 1.817 \times 10^{-19} \text{ J}$$

$$\nu = \frac{E}{h} = \frac{1.817 \times 10^{-19} \text{ J}}{6.626 \times 10^{-34} \text{ J s}} = 2.74 \times 10^{14} \text{ s}^{-1}$$

$$\Delta E_{5-3} = E_5 - E_3 = \left( \frac{-2.18 \times 10^{-18} \text{ J}}{5^2} \right) - \left( \frac{-2.18 \times 10^{-18} \text{ J}}{3^2} \right) = 1.550 \times 10^{-19} \text{ J}$$

$$\nu = \frac{E}{h} = \frac{1.550 \times 10^{-19} \text{ J}}{6.626 \times 10^{-34} \text{ J s}} = 2.34 \times 10^{14} \text{ s}^{-1}$$

These photons lie in the IR, just short of the visible spectral region (see Figure 4-4).

**11. (1 point)**

**For the following sets of quantum numbers, determine which describe actual orbitals and which are non-existent. For each one that is non-existent, list the restriction that forbids it:**

	<b>n</b>	<b>l</b>	<b>m<sub>l</sub></b>	<b>m<sub>s</sub></b>
<b>(a)</b>	<b>5</b>	<b>3</b>	<b>-2</b>	<b>-1</b>
<b>(b)</b>	<b>5</b>	<b>3</b>	<b>-3</b>	<b>+1/2</b>
<b>(c)</b>	<b>3</b>	<b>3</b>	<b>-3</b>	<b>+1/2</b>
<b>(d)</b>	<b>3</b>	<b>0</b>	<b>0</b>	<b>-1/2</b>

Remember that  $n$  must be a positive integer,  $l$  is restricted to zero and positive integers less than  $n$ ,  $m_l$  is

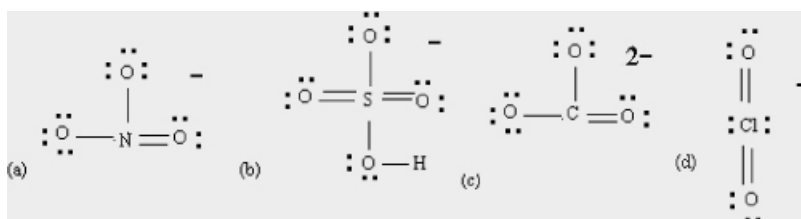
restricted to integers between  $-l$  and  $+l$ , and  $m_s = +\frac{1}{2}$  or  $-\frac{1}{2}$ .

(a) Non-existent:  $m_s$  must be  $+\frac{1}{2}$  or  $-\frac{1}{2}$ ; (b) actual; (c) non-existent:  $l$  must be less than  $n$ ; and (d) actual.

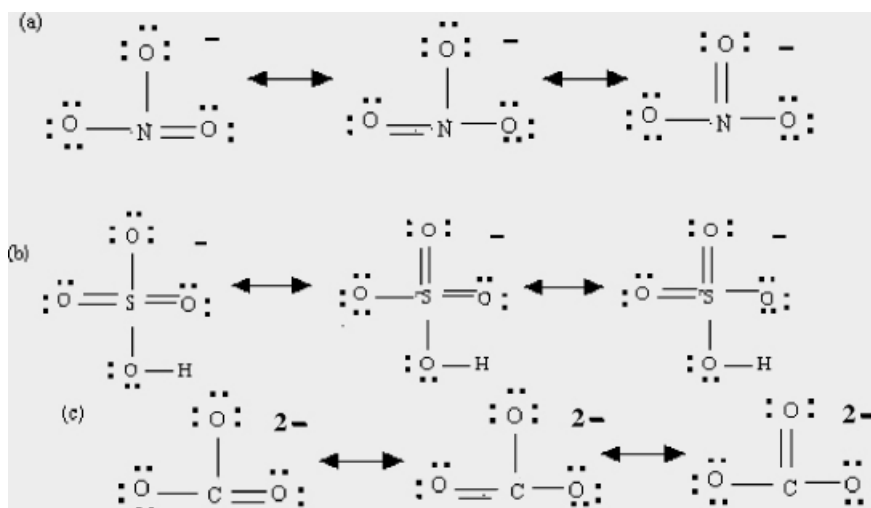
## 12. (2 points)

Determine the Lewis structure of each of the following polyatomic ions. Include all resonance structures and formal charges, where appropriate: (a)  $\text{NO}_3^-$ ; (b)  $\text{HSO}_4^-$ ; (c)  $\text{CO}_3^{2-}$ ; and (d)  $\text{ClO}_2^-$ .

. Optimize electron configurations of the inner atoms. The inner atoms in (a) and (c) are second row, so complete their octets. The inner atoms in (b) and (d) are third row, so reduce their formal charges to zero by making two double bonds to each:

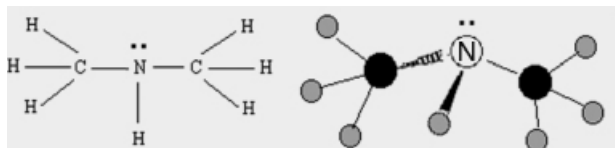


. The outer oxygen atoms all are equivalent, so all but (d) have three equivalent structures.



13. (1 points)

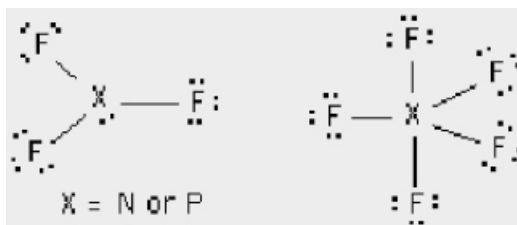
Write the Lewis structure of dimethylamine ((CH<sub>3</sub>)<sub>2</sub>NH). Determine the geometry. You can try to draw a ball-and-stick model of the molecule, showing the geometrical arrangement)



14. (1 point)

Both PF<sub>3</sub> and PF<sub>5</sub> are known compounds. NF<sub>3</sub> also exists, but NF<sub>5</sub> does not. Why is there no molecule with the formula NF<sub>5</sub>?

The Lewis structures of molecules with formula XF<sub>3</sub> show octets around the inner atom and FC<sub>X</sub> = 0, making them stable. Compounds with formula XF<sub>5</sub> also have FC<sub>X</sub> = 0 but have five electron pairs associated with the inner atom. This is possible for phosphorus, a third-row element that has *d* orbitals available for bonding. It is not possible for nitrogen, a second-row element that lacks valence *d* orbitals:



15. (1 point)

Identify the hybrid orbitals used by the inner atoms for:  
SO<sub>3</sub>, SO<sub>2</sub>, CHCl<sub>3</sub>, PBr<sub>3</sub>

. Use the Lewis structure of the molecule: (a) *sp*<sup>2</sup> hybrids; (b) *sp*<sup>2</sup> hybrids; (c) *sp*<sup>3</sup> *d* hybrids (d) *sp*<sup>3</sup> hybrids.

16 (3 points)

The condensation reaction of butadiene (C<sub>4</sub>H<sub>6</sub>) is second order in C<sub>4</sub>H<sub>6</sub>, with a rate constant of 0.93 M<sup>-1</sup> min<sup>-1</sup>. If the initial concentration of C<sub>4</sub>H<sub>6</sub> is 0.240 M, find (a) the

**time at which the concentration will be 0.100 M; and (b) the concentration after 25 min of reaction.**

This is stated to be a second-order reaction, so rate =  $k[\text{C}_4\text{H}_6]^2$  and applies:  $\frac{1}{[\text{A}]} - \frac{1}{[\text{A}]_0} = kt$

The problem states that  $k = 0.93 \text{ M}^{-1} \text{ min}^{-1}$ .

$$(a) kt = \frac{1}{0.100 \text{ M}} - \frac{1}{0.240 \text{ M}} = 5.83 \text{ M}^{-1}$$

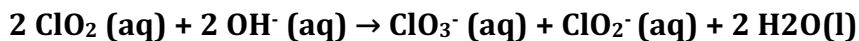
$$t = \frac{5.83 \text{ M}^{-1}}{0.93 \text{ M}^{-1} \text{ min}^{-1}} = 6.3 \text{ min}$$

$$(b) \frac{1}{[\text{A}]} = \left( \frac{1}{0.240 \text{ M}} \right) + (0.93 \text{ M}^{-1} \text{ min}^{-1})(25 \text{ min}) = 4.17 \text{ M}^{-1} + 23.25 \text{ M}^{-1} = 27.42 \text{ M}^{-1}$$

$$[\text{A}] = 3.6 \times 10^{-2} \text{ M}$$

**17. (4 points)**

**The following initial rate information was collected at 25 °C for this aqueous reaction:**



$[\text{ClO}_2]_0$ (M)	$[\text{OH}^-]_0$ (M)	Initial Rate (M/s)
0.050	0.050	$2.9 \times 10^{-2}$
0.075	0.050	$6.5 \times 10^{-2}$
0.075	0.075	$9.8 \times 10^{-2}$

**Determine the rate law and evaluate the rate constant for this reaction.**

From the data provided, we recognize this as an initial rate problem. The essential feature of the initial rate method is that we can take ratios of initial rates under different conditions. First, apply this technique to Experiments 1 and 2, which have the same initial concentration of  $\text{OH}^-$ :

$$\text{Initial rate}_1 = 2.9 \times 10^{-2} \text{ M/s} = k (0.050 \text{ M})^x (0.050 \text{ M})^y$$

$$\text{Initial rate}_2 = 6.5 \times 10^{-2} \text{ M/s} = k (0.075 \text{ M})^x (0.050 \text{ M})^y$$

When we take the ratio of the second initial rate to the first, the rate constant and the initial concentration term for OH<sup>-</sup> cancel:

$$\frac{\text{Initial rate}_2}{\text{Initial rate}_1} = \frac{6.5 \times 10^{-2} \text{ M/s}}{2.9 \times 10^{-2} \text{ M/s}} = \frac{k (0.075 \text{ M})^x (0.050 \text{ M})^y}{k (0.050 \text{ M})^x (0.050 \text{ M})^y} = \frac{(0.075 \text{ M})^x}{(0.050 \text{ M})^x}$$

Evaluating the ratios gives  $2.24 = (1.5)^x$ , from which  $x = 2$ .

Now repeat this analysis for the third experiment and the second experiment, for which the initial concentrations of ClO<sub>2</sub> are the same:

$$\frac{\text{Initial rate}_3}{\text{Initial rate}_2} = \frac{9.8 \times 10^{-2} \text{ M/s}}{6.5 \times 10^{-2} \text{ M/s}} = \frac{k (0.075 \text{ M})^x (0.075 \text{ M})^y}{k (0.075 \text{ M})^x (0.050 \text{ M})^y} = \frac{(0.075 \text{ M})^y}{(0.050 \text{ M})^y}$$

$1.5 = (1.5)^y$ , from which  $y = 1$ .

The rate law is as follows: Rate =  $k[\text{ClO}_2]^2[\text{OH}^-]$

$$1.5 = (1.5)^y$$

Use any of the experiments to evaluate the rate constant  $k$ :

$$2.9 \times 10^{-2} \text{ M/s} = k (0.050 \text{ M})^2 (0.050 \text{ M}) \quad k = \frac{2.9 \times 10^{-2} \text{ M/s}}{(0.050 \text{ M})^3} = 2.3 \times 10^2 \text{ M}^{-2} \text{ s}^{-1}$$

### 18. (2 points)

**The equilibrium constant for the dissociation of Cl<sub>2</sub> into atomic chlorine at 1200 K is  $2.5 \times 10^{-5}$ . If Cl<sub>2</sub> gas at 298 K, 0.57 bar is placed in a sealed container that is then heated to 1200 K, what is the equilibrium pressure of Cl atoms?**

To calculate concentrations at equilibrium from initial conditions, set up a concentration table. For a gas-phase reaction, concentrations must be expressed in bar. We know the pressure of Cl<sub>2</sub> gas at 298 K, from which we can calculate the initial pressure at 1200 K:

$$\frac{p_1}{T_1} = \frac{p_2}{T_2} \quad p_2 = \frac{p_1 T_2}{T_1} = \frac{(0.57 \text{ bar})(1200 \text{ K})}{(298 \text{ K})} = 2.3 \text{ bar}$$

Let  $-x$  = change in  $p_{\text{Cl}_2}$ :

Reaction:	Cl <sub>2</sub> (g)	2Cl (g)
Initial pressure (bar)	2.3	0
Change in pressure (bar)	-x	+2x
Equilibrium pressure (bar)	2.3 - x	2x

Now substitute into the equilibrium constant expression and solve for  $x$ :

$$K_{\text{eq}} = \frac{(p_{\text{Cl}})_{\text{eq}}^2}{(p_{\text{Cl}_2})_{\text{eq}}} = \frac{(2x)^2}{(2.3-x)} = 2.5 \times 10^{-5}; \text{ assume } x \ll 2.3$$

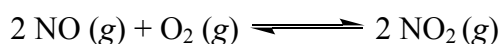
$$4x^2 = (2.5 \times 10^{-5})(2.3) = 5.75 \times 10^{-5} \quad x^2 = 1.44 \times 10^{-5}$$

$$x = 3.8 \times 10^{-3} \quad (p_{\text{Cl}})_{\text{eq}} = 3.8 \times 10^{-3} \text{ bar} \quad (p_{\text{Cl}_2})_{\text{eq}} = 2.3 \text{ bar}$$

0.0038  $\ll$  2.3, so the approximation is valid.

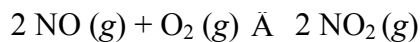
**19. (3 points)**

**Predict the effect of each of the following changes on the equilibrium position of the reaction:**



**(a) The partial pressure of NO<sub>2</sub> is cut in half. (b) The volume of the reactor is doubled. (c) A total of 10.0 bar of Ar gas is added to the reactor.**

To predict effects of changes on equilibrium position, apply Le Châtelier's principle: the system will respond in the direction that reduces the effect of the change. The reaction in Example 14-14 is

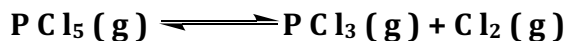


- (a) NO<sub>2</sub> is a product, so reducing its pressure causes the reaction to proceed to the right.
- (b) There are more moles of gas on the reactant side, so doubling the volume causes the reaction to proceed to the left.
- (c) Adding Ar does not change any of the pressures in the equilibrium expression, so this change has no effect.

effect.

**20 (3 points)**

**For the following reaction,  $K_{eq}$  is  $1.83 \times 10^{-3}$  at 390. K:**



**If 2.00 g of  $\text{PCl}_5$  is placed in a 3.00 L bulb at 390 K, what is the equilibrium pressure of  $\text{Cl}_2$**

To calculate concentrations at equilibrium from initial conditions, set up a concentration table. For a gas-phase reaction, concentrations must be expressed in bar. Use the ideal gas equation to calculate the initial pressure of  $\text{PCl}_5$  gas:

$$p = \frac{nRT}{V} = \frac{mRT}{MV} = \frac{(2.00 \text{ g})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(390 \text{ K})}{(208.22 \text{ g/mol})(3.00 \text{ L})} = 0.1025 \text{ atm}$$

Convert from atm to bar:

$$0.1025 \text{ atm} \left( \frac{1.013 \text{ bar}}{1 \text{ atm}} \right) = 0.1038 \text{ bar}$$

Let  $x$  = change in  $p_{\text{Cl}_2}$ :

Reaction:	$\text{PCl}_5$ (g)	$\text{PCl}_3$ (g) +	$\text{Cl}_2$ (g)
Initial pressure (bar)	0.1038	0	0
Change in pressure (bar)	$-x$	$+x$	$+x$
Equilibrium pressure (bar)	0.1038 $-x$	$x$	$x$

Now substitute into the equilibrium constant expression and solve for  $x$ . *Given the value of  $K$  there will be no assumption allowed in this case.* Thus :

$$K_{\text{eq}} = \frac{(p_{\text{PCl}_3})_{\text{eq}}(p_{\text{Cl}_2})_{\text{eq}}}{(p_{\text{PCl}_5})_{\text{eq}}} = \frac{(x)^2}{(0.1038-x)} = 1.83 \times 10^{-3}$$

$$x^2 = (1.83 \times 10^{-3})(0.1038-x) = (1.90 \times 10^{-4}) - (1.83 \times 10^{-3})x$$

$$x^2 + (1.83 \times 10^{-3})x - (1.90 \times 10^{-4}) = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-(1.83 \times 10^{-3}) \pm \sqrt{(1.83 \times 10^{-3})^2 + 4(1.90 \times 10^{-4})}}{2}$$

$$x = \frac{-(1.83 \times 10^{-3}) \pm \sqrt{(7.63 \times 10^{-4})}}{2} = \frac{-(1.83 \times 10^{-3}) + (2.76 \times 10^{-2})}{2} = 1.29 \times 10^{-2}$$

$$(p_{\text{PCl}_3})_{\text{eq}} = (p_{\text{Cl}_2})_{\text{eq}} = 1.29 \times 10^{-2} \text{ bar} \quad (p_{\text{PCl}_5})_{\text{eq}} = 0.1038 - 0.0129 = 0.0909 \text{ bar}$$

## 21. (2 points)

**Boric acid ( $\text{H}_3\text{BO}_3$ ),  $K_a = 5.4 \times 10^{-10}$ , is frequently used as an eyewash.) Calculate the pH of 0.050 M boric acid solution.**

(b) To calculate the pH of a solution, follow the standard procedure for equilibrium calculations:

Reaction: $\text{H}_2\text{O} +$	$\text{H}_3\text{BO}_3$	$\text{H}_2\text{B}^-$	$\text{H}^+$
	3	O	
		3	
		-	
		+	
Initial concentration (M)	0.050	0	0
Change in concentration (M)	-x	+x	+x
Final concentration (M)	0.050	x	x
	-		
	x		

$$K_a = 5.4 \times 10^{-10} = \frac{[\text{H}_2\text{BO}_3^-]_{\text{eq}}[\text{H}_3\text{O}^+]_{\text{eq}}}{[\text{H}_3\text{BO}_3]_{\text{eq}}} = \frac{x^2}{0.050-x}; \text{ assume that } x \ll 0.050:$$

$$5.4 \times 10^{-10} = \frac{x^2}{0.050}$$

$$x^2 = 2.7 \times 10^{-11}, \text{ so } x = 5.2 \times 10^{-6}; \text{ the assumption is valid.}$$

$$[\text{H}_3\text{O}^+] = 5.2 \times 10^{-6} \text{ M} \quad \text{pH} = -\log(5.2 \times 10^{-6}) = 5.28$$

**22. (2 points)**

**Vinegar is a dilute aqueous solution of acetic acid. A sample of vinegar has a pH of 2.39 and a density of 1.07 g/mL. What is the mass percentage of acetic acid in the vinegar? (acetic acid  $K_a = 1.8 \times 10^{-5}$ )**

To determine the mass percent of vinegar, we must first determine how much of each species is present in the solution using the standard procedure:

1. This is a weak acid. The major species are  $H_2O$  and  $CH_3CO_2H$ .

2. The dominant acid–base equilibrium is



3.  $K_a = 1.8 \times 10^{-5}$

4. In this solution, the pH corresponds to the equilibrium concentrations of hydronium ions and acetate ions:

$$[H_3O^+] = [CH_3CO_2^-] = 10^{-2.39} = 0.00407 \text{ M}$$

Assuming that the concentration of  $CH_3CO_2H$  is much larger than 0.00407 M:

$$K_a = 1.8 \times 10^{-5} = \frac{[H_3O^+][CH_3CO_2^-]}{[CH_3CO_2H]} = \frac{(0.00407)^2}{x}$$

$x = 0.92 \text{ M} = [CH_3CO_2H]$ ; the assumption is valid.

$$m_{\text{acetic acid}}/\text{L} = (0.92 \text{ mol})(60.05 \text{ g/mol}) = 55 \text{ g/L or } 0.055 \text{ g/mL acetic acid}$$

$$\text{mass}\% = 100\% \left( \frac{m_{\text{acetic acid}}/\text{mL}}{\rho_{\text{sample}}} \right) = 100\% \left( \frac{0.055 \text{ g/mL}}{1.07 \text{ g/mL}} \right) = 5.1\%$$

**23. (3 points)**

**What mass of sodium acetate must be added to 2.50 L of 0.55 M acetic acid to make a buffer solution of pH = 5.75? (acetic acid  $K_a = 1.8 \times 10^{-5}$ )**

This is a  $C_2H_3O_2^-/C_2H_3O_2H$  buffer. Use the buffer equation, Equation 16-1:  $M_{NaC_2H_3O_2} = 22.99 \text{ g/mol} + 2(12.01 \text{ g/mol}) + 3(1.01 \text{ g/mol}) + 2(16.00 \text{ g/mol}) = 82.04 \text{ g/mol}$

$$\log\left(\frac{n_{\text{C}_2\text{H}_3\text{O}_2^-}}{n_{\text{C}_2\text{H}_3\text{O}_2\text{H}}}\right) = \text{pH} - \text{p}K_a = 5.75 - 4.75 = 1.00$$

$$\frac{n_{\text{C}_2\text{H}_3\text{O}_2^-}}{n_{\text{C}_2\text{H}_3\text{O}_2\text{H}}} = 10^{1.00}$$

$$n_{\text{C}_2\text{H}_3\text{O}_2^-} = 10.0(\text{mol C}_2\text{H}_3\text{O}_2\text{H}) = 10.0(0.55 \text{ M})(2.50 \text{ L}) = 13.8 \text{ moles of NaC}_2\text{H}_3\text{O}_2$$

$$m_{\text{NaC}_2\text{H}_3\text{O}_2} = (13.8 \text{ mol})(82.04 \text{ g/mol}) = 1.1 \times 10^3 \text{ g}$$

#### 24. (5 points)

A sample of carbonic acid (0.125 L, 0.120 M,  $\text{p}K_{a1} = 6.35$ ,  $\text{p}K_{a2} = 10.33$ ) was titrated with 1.504 M NaOH. Calculate the pH at the following points: (a) before the titration; (b) the first midpoint; (c) the first stoichiometric point; (d) the second midpoint; and (e) the second stoichiometric point.

The major species present are different at various points during a titration, so there are different dominant equilibria that must be identified before doing an equilibrium calculation to determine pH.

(a) Before titration begins, the major species are a weak acid, carbonic acid (HA) and  $\text{H}_2\text{O}$ , and the dominant equilibrium is  $\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{A}^- + \text{H}_3\text{O}^+$ .

$$K_{a1} = 4.5 \times 10^{-7} \text{ and } \text{p}K_{a1} = 6.35:$$

Reaction: $\text{H}_2\text{O} +$	HA	$\text{A}^-$ +	$\text{H}_3\text{O}^+$
Initial concentration (M)	0.120	0	0
Change in concentration (M)	$-x$	$+x$	$+x$
Final concentration (M)	$0.120 - x$	$x$	$x$

Now substitute into the equilibrium constant expression and solve for  $x$ :

$$K_{a1} = 4.5 \times 10^{-7} = \frac{[A^-]_{\text{eq}}[H_3O^+]_{\text{eq}}}{[HA]_{\text{eq}}} = \frac{x^2}{0.120 - x}; \text{ assume that } x \ll 0.120:$$

$$x^2 = (4.5 \times 10^{-7})(0.120) = 5.4 \times 10^{-8}, \text{ so } x = 2.3 \times 10^{-4}; \text{ the assumption is valid.}$$

$$[H_3O^+] = 2.3 \times 10^{-4} \text{ M} \quad \text{pH} = -\log(2.3 \times 10^{-4}) = 3.63$$

(b) At the first midpoint, both acid and conjugate base of carbonic acid are present in equal concentrations, the solution is buffered, and the pH can be calculated using the buffer equation:

$$\text{pH} = \text{p}K_{a1} + \log 1 = 6.35$$

(c) At the first stoichiometric point, Equation 16-2 from your textbook applies:

$$[H_3O^+]_{\text{eq, 1st stoichiometric point}} = \sqrt{K_{a1}K_{a2}}$$

$$[H_3O^+] = \sqrt{(4.5 \times 10^{-7})(4.7 \times 10^{-11})} = \sqrt{(21.1 \times 10^{-18})} = 4.6 \times 10^{-9} \text{ M}$$

$$\text{pH} = -\log(4.6 \times 10^{-9}) = 8.34$$

(d) At the second midpoint, both acid and conjugate base of hydrogen carbonate are present in equal concentrations, the solution is buffered, and the pH can be calculated using the buffer equation:

$$\text{pH} = \text{p}K_{a2} + \log 1 = 10.33$$

(e) At the second stoichiometric point, all the hydrogen carbonate has been converted into its conjugate base, so the major acid-base species present are carbonate ( $A^{2-}$ ) and  $H_2O$  and the dominant equilibrium is



$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.7 \times 10^{-11}} = 2.1 \times 10^{-4}$$

Reaction: $H_2O +$	$A^-$	$HA^-$ +	$OH^-$
Initial concentration (M)	0.120	0	0

Change in concentration (M)	$-x$	$+x$	$+x$
Final concentration (M)	$0.120 - x$	$x$	$x$

Now substitute into the equilibrium constant expression and solve for  $x$ :

$$K_b = 2.1 \times 10^{-4} = \frac{[\text{HA}]_{\text{eq}}[\text{OH}^-]_{\text{eq}}}{[\text{A}^-]_{\text{eq}}} = \frac{x^2}{0.120 - x}$$

$$x^2 = (2.1 \times 10^{-4})(0.120 - x) = (2.55 \times 10^{-5}) - (2.1 \times 10^{-4})x$$

$$x^2 + (2.1 \times 10^{-4})x - (2.55 \times 10^{-5}) = 0$$

$[\text{OH}^-]_{\text{eq}} =$

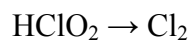
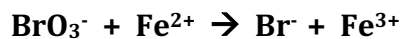
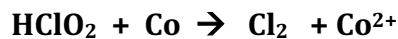
$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-(2.1 \times 10^{-4}) \pm \sqrt{(2.1 \times 10^{-4})^2 - 4(-2.55 \times 10^{-5})}}{2} = 4.9 \times 10^{-3}$$

$$4.9 \times 10^{-3} \text{ M}$$

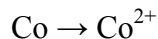
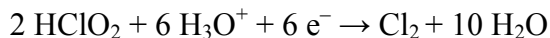
$$\text{pOH} = -\log(4.9 \times 10^{-3}) = 2.31 \quad \text{pH} = 14.00 - \text{pOH} = 14.00 - 2.31 = 11.69$$

25. (2 points)

Balance the following redox reaction in acidic conditions:

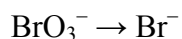
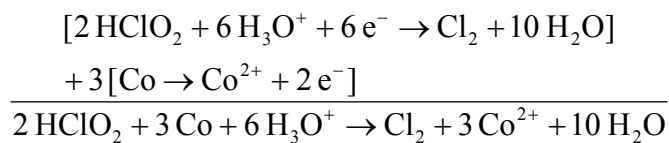


(1) Multiply  $\text{HClO}_2$  by 2; (2) add 4  $\text{H}_2\text{O}$  on the right; (3) add 6  $\text{H}_3\text{O}^+$  on the left and 6  $\text{H}_2\text{O}$  on the right; (4) add 6  $\text{e}^-$  on the left:

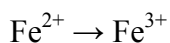


(1–3) All elements are already balanced; (4) add 2  $\text{e}^-$  on right:  $\text{Co} \rightarrow \text{Co}^{2+} + 2 \text{e}^-$

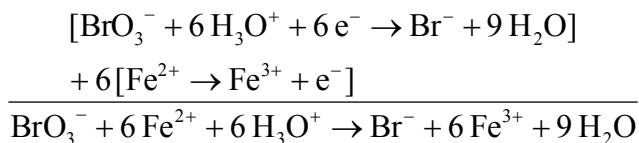
Multiply the second reaction by 3 and add:

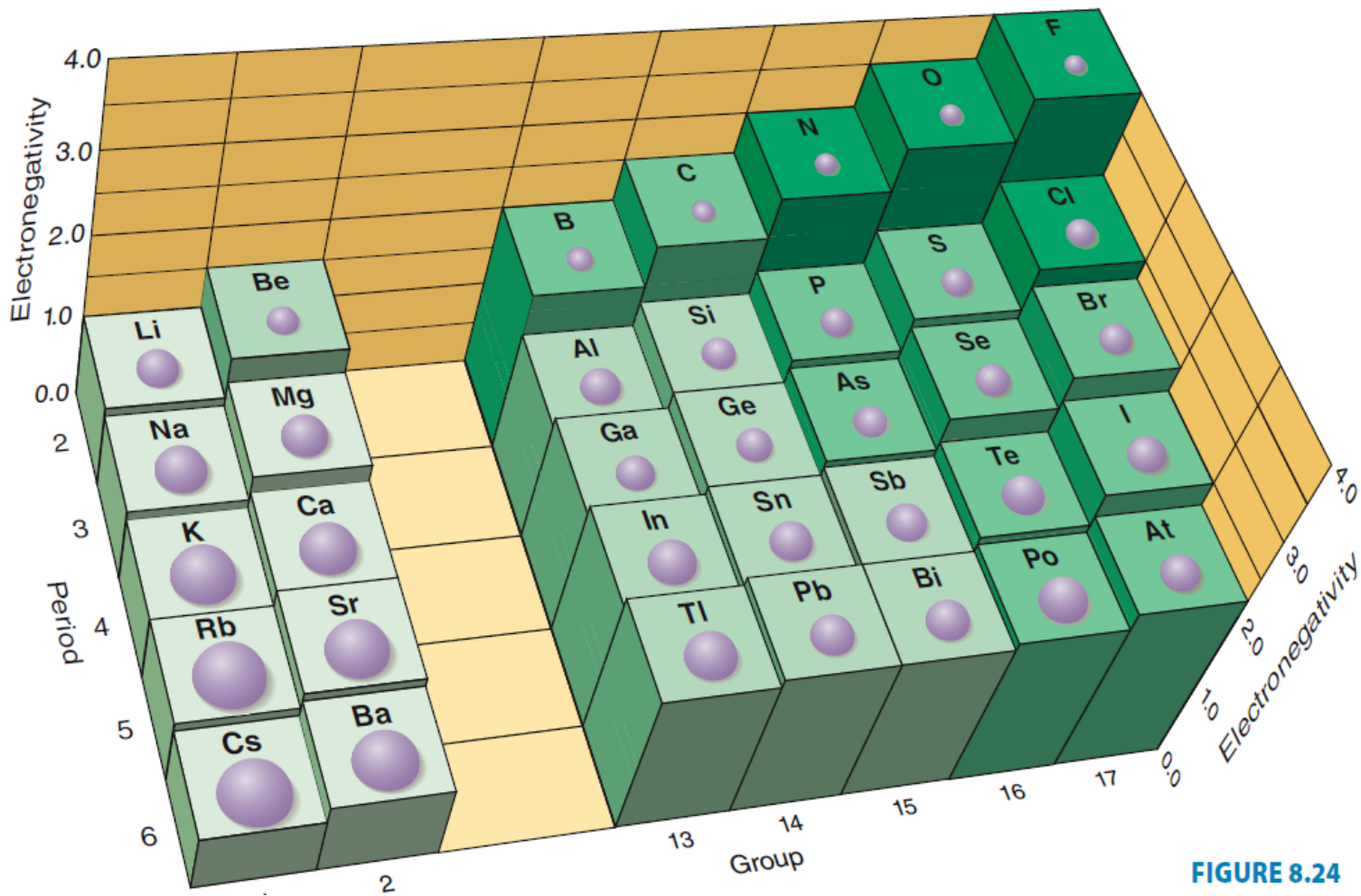


(1) Br is balanced; (2) add 3  $\text{H}_2\text{O}$  on the right; (3) add 6  $\text{H}_3\text{O}^+$  on the left and 6  $\text{H}_2\text{O}$  on the right; (4) add 6  $\text{e}^-$  on the left:  $\text{BrO}_3^- + 6 \text{H}_3\text{O}^+ + 6 \text{e}^- \rightarrow \text{Br}^- + 9 \text{H}_2\text{O}$



(1–3) All elements are balanced; (4) add 1  $\text{e}^-$  on the right:  $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$  Multiply the second reaction by 6 and add:



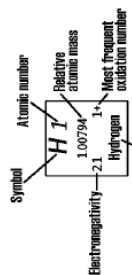


**FIGURE 8.24**

# Mokleur's Periodic table of the elements

18  
VIIIA

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	
1A	2A	3A	4A	5A	6A	7A	8	VIII	10	11B	12	13A	14A	15A	16A	17A	VIIIA	
<b>H</b> 1 1.00794 2.1 Hydrogen	<b>He</b> 2 4.002602 Helium	<b>Li</b> 3 6.941 1.0 1.5 2+ Lithium	<b>Be</b> 4 9.012182 1.0 1.5 2+ Beryllium	<b>B</b> 5 10.811 2.0 3+ Boron	<b>C</b> 6 12.011 2.5 4+ Carbon	<b>N</b> 7 14.00674 3.0 3+ Nitrogen	<b>O</b> 8 15.9994 3.5 2- Oxygen	<b>F</b> 9 18.9984032 4.0 1- Fluorine	<b>Ne</b> 10 20.1797 Helium	<b>Na</b> 11 22.989768 0.9 1+ Sodium	<b>Mg</b> 12 24.30509 1.2 2+ Magnesium	<b>Al</b> 13 26.981539 1.3 3+ Aluminum	<b>Si</b> 14 28.0855 1.8 4+ Silicon	<b>P</b> 15 30.973762 2.1 3+ Phosphorus	<b>S</b> 16 32.066 1.8 4+ Sulfur	<b>Cl</b> 17 35.4527 3.0 1- Chlorine	<b>Ar</b> 18 39.948 Argon	
<b>K</b> 19 39.0983 0.8 1+ Potassium	<b>Ca</b> 20 40.078 1.0 2+ Calcium	<b>Sc</b> 21 44.955910 1.3 3+ Scandium	<b>Ti</b> 22 47.88 1.5 3+ Titanium	<b>V</b> 23 50.9415 1.6 3+ Vanadium	<b>Cr</b> 24 51.9961 1.6 3+ Chromium	<b>Mn</b> 25 54.93805 1.5 2+ Manganese	<b>Fe</b> 26 55.847 1.8 2+ Iron	<b>Co</b> 27 58.9332 1.8 2+ Cobalt	<b>Ni</b> 28 58.6934 1.8 2+ Nickel	<b>Cu</b> 29 63.546 1.9 2+ Copper	<b>Zn</b> 30 65.39 1.6 2+ Zinc	<b>Ga</b> 31 69.723 1.6 3+ Gallium	<b>Ge</b> 32 72.61 2.0 4+ Germanium	<b>As</b> 33 74.92159 2.0 3+ Arsenic	<b>Se</b> 34 78.96 2.4 2- Selenium	<b>Br</b> 35 79.904 2.8 1- Bromine	<b>Kr</b> 36 83.80 Krypton	
<b>Rb</b> 37 85.4678 0.8 1+ Rubidium	<b>Sr</b> 38 87.62 1.3 2+ Strontium	<b>Y</b> 39 88.90585 1.3 3+ Yttrium	<b>Zr</b> 40 91.224 1.4 4+ Zirconium	<b>Nb</b> 41 92.90638 1.6 5+ Niobium	<b>Mo</b> 42 95.94 1.8 6+ Molybdenum	<b>Tc</b> 43 98.9063 1.9 7+ Technetium	<b>Ru</b> 44 101.57 2.2 3+ Ruthenium	<b>Rh</b> 45 102.9055 2.2 3+ Rhodium	<b>Pd</b> 46 106.42 2.2 2+ Palladium	<b>Ag</b> 47 107.8682 1.9 1+ Silver	<b>Cd</b> 48 112.411 1.7 2+ Cadmium	<b>In</b> 49 114.82 1.7 3+ Indium	<b>Sn</b> 50 118.71 2.0 4+ Tin	<b>Sb</b> 51 121.757 1.9 3+ Antimony	<b>Te</b> 52 127.60 2.1 4+ Tellurium	<b>I</b> 53 126.90447 2.5 1- Iodine	<b>Xe</b> 54 131.29 Xenon	
<b>Cs</b> 55 132.90543 0.7 1+ Cesium	<b>Ba</b> 56 137.327 1.1 2+ Barium	<b>La</b> 57 138.9055 1.1 3+ Lanthanum	<b>Hf</b> 72 178.49 1.3 4+ Hafnium	<b>Ta</b> 73 180.9479 1.5 5+ Tantalum	<b>W</b> 74 183.85 1.7 6+ Tungsten	<b>Re</b> 75 186.207 1.9 7+ Rhenium	<b>Os</b> 76 190.2 2.2 4+ Osmium	<b>Ir</b> 77 192.22 2.2 4+ Iridium	<b>Pt</b> 78 195.08 2.2 2+ Platinum	<b>Au</b> 79 196.96654 2.0 3+ Gold	<b>Hg</b> 80 200.59 1.9 2+ Mercury	<b>Tl</b> 81 204.3833 1.8 1+ Thallium	<b>Pb</b> 82 207.2 2.4 2+ Lead	<b>Bi</b> 83 208.98037 1.9 3+ Bismuth	<b>Po</b> 84 209 2.0 2- Polonium	<b>At</b> 85 209 2.2 - Astatine	<b>Rn</b> 86 222.0176 Radon	
<b>Fr</b> 87 223.0197 0.9 1+ Francium	<b>Ra</b> 88 226.0254 0.9 2+ Radium	<b>Ac</b> 89 227.0278 1.1 3+ Actinium	<b>Rf</b> 104 261.11 1.3 4+ Rutherfordium	<b>Db</b> 105 262.11 1.4 5+ Dubnium	<b>Sg</b> 106 263.12 1.3 6+ Seaborgium	<b>Bh</b> 107 262.12 1.3 7+ Bohrium	<b>Hs</b> 108 264 1.3 8+ Hassium	<b>Mt</b> 109 266.1378 1.3 9+ Meitnerium	<b>Uun</b> 110 269 1.3 10+ Ununium	<b>Uuu</b> 111 272 1.3 11+ Ununium	<b>Uu</b> 112 277 1.3 12+ Ununium	<b>Uuq</b> 114 289 1.3 14+ Unquadium	<b>Uuh</b> 116 289 1.3 16+ Unhexium	<b>Uub</b> 118 293 1.3 18+ Unoctium				



6	<b>Ce</b> 58 140.115 1.1 3+ Cerium	<b>Pr</b> 59 140.90765 1.1 3+ Praseodymium	<b>Nd</b> 60 144.24 1.1 3+ Neodymium	<b>Pm</b> 61 144.9127 1.1 3+ Promethium	<b>Sm</b> 62 150.36 1.2 3+ Samarium	<b>Eu</b> 63 151.965 1.2 3+ Europium	<b>Gd</b> 64 157.25 1.2 3+ Gadolinium	<b>Tb</b> 65 168.93032 1.2 3+ Terbium	<b>Dy</b> 66 162.50 1.2 3+ Dysprosium	<b>Ho</b> 67 164.93032 1.2 3+ Holmium	<b>Er</b> 68 167.26 1.2 3+ Erbium	<b>Tm</b> 69 168.93421 1.2 3+ Thulium	<b>Yb</b> 70 173.04 1.1 3+ Ytterbium	<b>Lu</b> 71 174.967 1.2 3+ Lutetium
7	<b>Th</b> 90 232.0381 1.3 4+ Thorium	<b>Pa</b> 91 231.03588 1.5 5+ Protactinium	<b>U</b> 92 238.02891 1.4 6+ Uranium	<b>Np</b> 93 237.04817 1.3 5+ Neptunium	<b>Pu</b> 94 244.06422 1.3 6+ Plutonium	<b>Am</b> 95 243.06115 1.3 5+ Americium	<b>Cm</b> 96 247 1.3 6+ Curium	<b>Bk</b> 97 247.0703 1.3 7+ Berkelium	<b>Cf</b> 98 251.0796 1.3 8+ Californium	<b>Es</b> 99 252.083 1.3 9+ Einsteinium	<b>Fm</b> 100 257.0951 1.3 10+ Fermium	<b>Md</b> 101 258.10 1.3 11+ Mendelevium	<b>No</b> 102 259.1009 1.3 12+ Nobelium	<b>Lr</b> 103 260.1053 1.3 13+ Lawrencium

Under normal conditions, bold symbols correspond to solid state, bold italic correspond to liquid state, italic correspond to gaseous state and normal correspond to synthetic elements.

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## Data For Water

Density = 1.00 g/mL (at 25°C)

$s = 2.13 \text{ J g}^{-1} \text{ K}^{-1}$  (solid)

$s = 4.184 \text{ J g}^{-1} \text{ K}^{-1}$  (liquid)

$s = 2.01 \text{ J g}^{-1} \text{ K}^{-1}$  (gas)

$\Delta H_{\text{fus}}^{\circ} = 6.02 \text{ kJ mol}^{-1}$

$\Delta H_{\text{vap}}^{\circ} = 40.7 \text{ kJ mol}^{-1}$

## Constants and Conversion Factors

1 mmHg = 1 torr    760 mmHg = 1 atm    1 atm = 101.325 kPa    1 atm = 1.013125 bar  
1 cm<sup>3</sup> = 1 mL    1000 mL = 1 L    1000 L = 1 m<sup>3</sup>

Avogadro's Number	$N$	$6.022 \times 10^{23} \text{ mol}^{-1}$
Boltzmann's constant	$k$	$1.30866 \times 10^{-23} \text{ J} \cdot \text{K}^{-1}$
Faraday's constant	$F$	$96,485 \text{ C} \cdot \text{mol}^{-1}$
Gas constant	$R$	$8.31451 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$
	$R$	$0.08206 \text{ atm} \cdot \text{L} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$
	$R$	$8.31451 \text{ m}^3 \text{ Pa} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$
	$R$	$0.0831451 \text{ bar} \cdot \text{L} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$
Planck's constant	$h$	$6.62608 \times 10^{-34} \text{ J} \cdot \text{s}$
Speed of Light	$c$	$2.99792458 \times 10^8 \text{ m} \cdot \text{s}^{-1}$

## Gas Laws

$$PV = nRT$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$P_T = P_1 + P_2 + P_3 + \dots$$

$$d = \frac{m}{V} = \frac{P \cdot MM}{RT}$$

$$E_K = \frac{1}{2}mv^2$$

$$u_{rms} = \sqrt{\frac{3RT}{MM}}$$

$$\frac{\text{Rate A}}{\text{Rate B}} = \sqrt{\frac{MM_B}{MM_A}}$$

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

## Equilibrium

$$K_p = K_c(RT)^{\Delta n}$$

## Acid/Base

$$pOH = -\log[OH^-]$$

$$pH = -\log[H^+]$$

$$pH + pOH = 14$$

$$K_a \times K_b = K_w$$

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$$pH = \frac{pK_{a1} + pK_{a2}}{2}$$

## Thermochemistry

$$\Delta U = q + W$$

$$W_{\text{system}} = -P\Delta V = -\Delta nRT$$

$$\Delta H = \Delta U + P\Delta V$$

$$q_p = \Delta U + P\Delta V$$

$$q = ms\Delta T$$

$$\Delta H_{\text{rxn}}^\circ = \sum n\Delta H_f^\circ(\text{pds}) - \sum n\Delta H_f^\circ(\text{rxts})$$

## The atom

$$E = hv$$

$$c = v\lambda$$

$$E = -B/n^2$$

## Kinetics

$$[A]_t = [A]_o - kt$$

$$\ln[A]_t = \ln[A]_o - kt$$

$$1/[A]_t = 1/[A]_o + kt$$

$$k = Ae^{(-E_a/RT)}$$

$$\ln(k_2/k_1) = (-E_a/R)(1/T_2 - 1/T_1)$$



