Final exam 17 December 2017, questions and answers

Principles of Chemistry (University of Ottawa)

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Final Exam – CHM 1311-F

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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.**

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

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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:

$$2\,NH_3(g) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)$$

What mass (in kg) of NH3 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: $2\ NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$. The result is asked for in kilograms; thus, the kg to g mass unit conversions will cancel and do not need to be shown:

$$3.50\,\text{ton}\,(NH_4)_2SO_4\left(\frac{10^3\,\text{kg}}{1\,\text{ton}}\right)\left(\frac{1\,\text{mol}}{132.1\,\text{g}}\right)\left(\frac{2\,\text{mol}\,NH_3}{1\,\text{mol}\,(NH_4)_2SO_4}\right)\left(\frac{17.04\,\text{g}}{1\text{ mol}}\right)=903\,\text{kg}\,NH_3$$

2. (1 point)
The chemical formula for aluminum sulphate is $Al_2(SO_4)_3$. (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) $n = \dfrac{m}{M} = 25.0\text{ g}\left(\dfrac{1\text{ mol}}{342.17\text{ g}}\right) = 7.31\times10^{-2}\text{ mol}$

(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):

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

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

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

(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_4) is an intermediate in the production of uranium hexafluoride (UF_6). One method of manufacturing UF4 consists of reacting UO_2 with HF according to the (unbalanced) reaction $UO_2 + HF \rightarrow UF_4 + H_2O$. Assuming a yield of 99%, calculate the volume of 31.8 M HF required to manufacture 25.0 tonnes of UF_4

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:

$UO_2 + 4\ HF \rightarrow UF_4 + H_2O$

Now the H and O atoms can be balanced by increasing the coefficient of H_2O to 2:

$UO_2 + 4\ HF \rightarrow UF_4 + 2\ H_2O$

25.0 tonnes of UF_4 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_4 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 × 10^3 mL contains chlorine gas at 27 °C and 1.25 × 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_2, a gas. The product is NaCl. The balanced chemical reaction is

$2 \text{ Na } (s) + \text{Cl}_2 (g) \rightarrow 2 \text{ NaCl } (s)$.

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:

$$p_{\mathrm{Cl_2}} = 1.25 \times 10^3 \text{ Torr}\left(\frac{1 \text{ bar}}{750.06 \text{ Torr}}\right) = 1.67 \text{ bar}$$

$$n_{\mathrm{Cl_2}} = \frac{pV}{RT} = \frac{(1.67 \text{ bar})(3.00 \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(300 \text{ K})} = 0.200 \text{ mol}$$

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

$$\mathrm{Cl_2}\text{: } 0.200 \text{ mol} \qquad \text{Na: } \frac{0.300 \text{ mol}}{2} = 0.\ 150 \text{ mol (LR)}$$

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

Reaction:	**2 Na (*s*) +**	**Cl$_2$ (*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$_2$ 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 $n = \frac{m}{M}$:

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

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 CF_2Cl_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 CO_2 and H_2O 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 CO_2, which the mouse exhales. The amount of oxygen consumed is therefore equal to the amount of CO_2 exhaled. The solid KOH in the chamber absorbs all the 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

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

The amount of oxygen consumed is therefore 0.08439 – 0.07998 mol = 0.00441 mol.

(0.00441 mol)(32.00 g/mol) = 0.141 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? (<u>molar</u> c_{H2O} = 75.291 J mol^{-1} °C^{-1}; <u>molar</u> c_{Fe} = 25.10 J mol^{-1} °C^{-1})

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

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

$$q_{\text{kettle}} = (24.2\text{ mol})(25.10\text{ J mol}^{-1}\ ^\circ\text{C}^{-1})(72.0\ ^\circ\text{C}) = 4.37\times10^{4}\text{ J}$$

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

$$q_{\text{water}} = (153\text{ mol})(75.291\text{ J mol}^{-1}\ ^\circ\text{C}^{-1})(72.0\ ^\circ\text{C}) = 8.29\times10^{5}\text{ J}$$

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

8. (2 points)

Acetylene (C_2H_2) is used in welding torches because it has a high enthalpy of combustion. When 1.00 g of acetylene burns completely in excess O_2 gas at constant volume, it releases 48.2 kJ of energy. (a) What is the balanced chemical equation i for this reaction? (b) What is the molar energy of combustion of acetylene? (c) How much energy is released per mole of O_2 consumed?

In a combustion reaction, the products are CO_2 and H_2O:

$$C_2H_2 + O_2 \rightarrow CO_2 + H_2O \text{ (unbalanced)}$$

Follow standard procedures to balance the equation. Give CO_2 a coefficient of 2 to balance C:

$$C_2H_2 + O_2 \rightarrow 2\ CO_2 + H_2O$$

$$2\ C + 2\ H + 2\ O \rightarrow 2\ C + 2\ H + 5\ O$$

Give O_2 a coefficient of 5/2 to balance O, and then multiply by 2 to clear fractions:

$$2\ C_2H_2 + 5\ O_2 \rightarrow 4\ CO_2 + 2\ H_2O$$

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

$$\Delta E = \left(\frac{-48.2\text{ kJ}}{1.00\text{ g}}\right)\left(\frac{26.04\text{ g}}{1\text{ mol}}\right) = -1.26\times10^{3}\text{ kJ/mol }C_2H_2$$

(c) Five moles of O_2 is consumed for every 2 moles of acetylene, so the energy released per mole of O_2 is

$$\Delta E = \left(\frac{-1.26\times10^{3}\text{ kJ}}{1\text{ mol acetylene}}\right)\left(\frac{2\text{ mol acetylene}}{5\text{ mol }O_2}\right) = -5.04\times10^{2}\text{ kJ/mol }O_2$$

9. (4 points)

When light of frequency $1.30 \times 10^{15}\ s^{-1}$ shines on the surface of cesium metal, electrons are ejected with a maximum kinetic energy of 5.2×10^{-19} 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 (a) $\lambda = \dfrac{c}{\nu} = \dfrac{2.998 \times 10^{8}\text{ m/s}}{1.30 \times 10^{15}\text{ s}^{-1}} = 2.31 \times 10^{-7}\text{ m}$

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

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

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:

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

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

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

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

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

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:

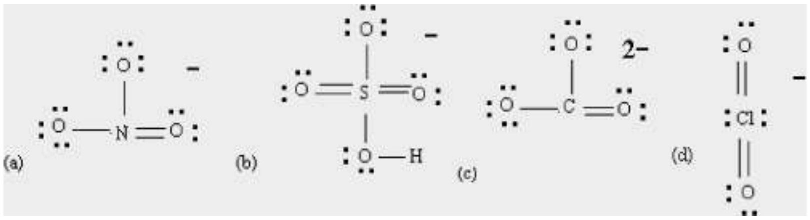
	n	**l**	**m_l**	**m_s**
(a)	**5**	**3**	**−2**	**−1**
(b)	**5**	**3**	**−3**	**+1/2**
(c)	**3**	**3**	**−3**	**+1/2**
(d)	**3**	**0**	**0**	**−1/ 2**

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 $-\frac{1}{2}$.

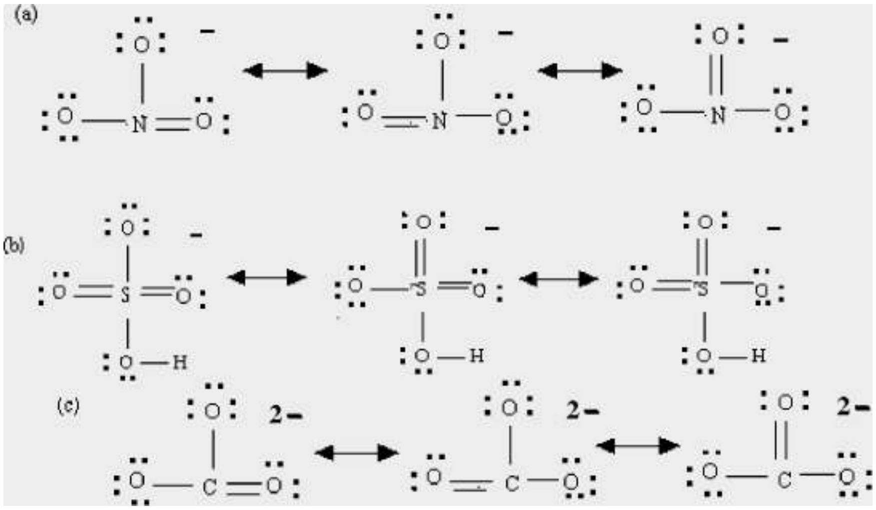
(a) Non-existent: m_s must be $+\frac{1}{2}$ or $-\frac{1}{2}$; (b) actual; (c) non-existent: l must be less than n; and (d) actual.

12. (2 points)
Determine the Lewis structure of each of the following polyatomic ions. Include all resonance structures and formal charges, where appropriate: (a) NO_3^-; (b) HSO_4^-; (c) CO_3^{2-}; and (d) 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:

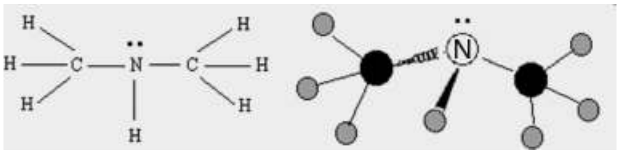


. The outer oxygen atoms all are equivalent, so all but (d) have three equivalent structures.



13. (1 points)

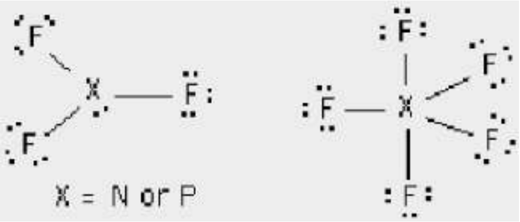
Write the Lewis structure of dimethylamine (($(CH_3)_2NH$). Determine the geometry. You can try to draw a ball-and-stick model of the molecule, showing the geometrical arrangement)



14. (1 point)

Both PF_3 and PF_5 are known compounds. NF_3 also exists, but NF_5 does not. Why is there no molecule with the formula NF_5?

The Lewis structures of molecules with formula XF_3 show octets around the inner atom and $FC_X = 0$, making them stable. Compounds with formula XF_5 also have $FC_X = 0$ 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_3, SO_2, $CHCl_3$, PBr_3

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

16 (3 points)

The condensation reaction of butadiene (C_4H_6) is second order in C_4H_6, with a rate constant of 0.93 M^{-1} min^{-1}. If the initial concentration of C_4H_6 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 = $k[C_4H_6]^2$ and applies: $\frac{1}{[A]} - \frac{1}{[A]_0} = kt$

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

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

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

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

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

17. (4 points)

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

$2\ ClO_2\ (aq) + 2\ OH^-\ (aq) \rightarrow ClO_3^-\ (aq) + ClO_2^-\ (aq) + 2\ H2O(l)$

$[ClO_2]_0$ (M)	**$[OH-]_0$ (M)**	**Initial Rate (M/s)**
0.050	**0.050**	**2.9 × 10-2**
0.075	**0.050**	**6.5 × 10-2**
0.075	**0.075**	**9.8 × 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 OH^-:

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

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

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

$$\frac{\text{Initial rate}_2}{\text{Initial rate}_1}=\frac{6.5\times10^{-2}\ \text{M/s}}{2.9\times10^{-2}\ \text{M/s}}=\frac{k\ (0.075\ \text{M})^x\ (0.050\ \text{M})^y}{k\ (0.050\ \text{M})^x\ (0.050\ \text{M})^y}=\frac{(0.075\ \text{M})^x}{(0.050\ \text{M})^x}$$

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_2 are the same:

$$\frac{\text{Initial rate}_3}{\text{Initial rate}_2}=\frac{9.8\times10^{-2}\ \text{M/s}}{6.5\times10^{-2}\ \text{M/s}}=\frac{k\ (0.075\ \text{M})^x\ (0.075\ \text{M})}{k\ (0.075\ \text{M})^x\ (0.050\ \text{M})^y}=\frac{(0.075\ \text{M})^x}{(0.050\ \text{M})^x}$$

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

The rate law is as follows: Rate = $k[ClO_2]^2[OH^-]$

$$1.5 = (1.5)^y$$

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

$$2.9\times10^{-2}\ \text{M/s} = k\ (0.050\ \text{M})^2(0.050\ \text{M}) \qquad k=\frac{2.9\times10^{-2}\ \text{M/s}}{(0.050\ \text{M})^3}=2.3\times10^{2}\ \text{M}^{-2}\ \text{s}^{-1}$$

18. (2 points)

The equilibrium constant for the dissociation of Cl_2 into atomic chlorine at 1200 K is 2.5×10^{-5}. If Cl_2 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_2 gas at 298 K, from which we can calculate the initial pressure at 1200 K:

$$\frac{p_1}{T_1}=\frac{p_2}{T_2} \qquad p_2=\frac{p_1T_2}{T_1}=\frac{(0.57\ \text{bar})\ (1200\ \text{K})}{(298\ \text{K})}=2.3\ \text{bar}$$

Let $-x$ = change in p_{Cl_2}:

Reaction:	Cl_2 (*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:

$$K_{eq} = \frac{(p_{Cl})^2_{eq}}{(p_{Cl_2})_{eq}} = \frac{(2x)^2}{(2.3 - x)} = 2.5 \text{ x } 10^{-5}\text{; assume } x << 2.3$$

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

$$x = 3.8 \times 10^{-3} \qquad (p_{Cl})_{eq} = 3.8 \times 10^{-3} \text{ bar } (p_{Cl_2})_{eq} = 2.3 \text{ bar}$$

0.0038 << 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:

$$2\ NO\ (g) + O_2\ (g) \rightleftharpoons 2\ NO_2\ (g)$$

(a) The partial pressure of NO_2 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

2 NO (*g*) + O_2 (*g*) Ä 2 NO_2 (*g*)

(a) NO_2 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.

20 (3 points)

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

$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$

If 2.00 g of PCl_5 is placed in a 3.00 L bulb at 390 K, what is the equilibrium pressure of 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 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_{Cl_2}:

Reaction:	PCl_5 (g)	PCl_3 (g) +	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 :

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

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

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

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

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

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

21. (2 points)
Boric acid (H_3BO_3), Ka = 5.4 × 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: H_2O +	H_3BO_3	$H_2BO_3^-$ +	+
Initial concentration (M)	0.050	0	0
Change in concentration (M)	$-x$	$+x$	$+x$
Final concentration (M)	$0.050 - x$	x	x

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

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

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

$[\text{H}_3\text{O}^+] = 5.2 \times 10^{-6}$ M 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

$$CH_3CO_2H + H_2O \rightleftharpoons CH_3CO_2^- + H_3O^+$$

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.

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

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

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 g/mol + 2(12.01 g/mol) + 3(1.01 g/mol) + 2(16.00 g/mol) = 82.04 g/mol

$$\log\left(\frac{n_{C_2H_3O_2^-}}{n_{C_2H_3O_2H}}\right) = pH - pK_a = 5.75 - 4.75 = 1.00$$

$$\frac{n_{C_2H_3O_2^-}}{n_{C_2H_3O_2H}} = 10^{1.00}$$

$$n_{C_2H_3O_2^-} = 10.0(\text{mol } C_2H_3O_2H) = 10.0(0.55\text{ M})(2.50\text{ L}) = 13.8\text{ moles of } NaC_2H_3O_2$$

$$m_{NaC_2H_3O_2} = (13.8\text{ mol})(82.04\text{ g/mol}) = 1.1 \times 10^3\text{ g}$$

24. (5 points)

A sample of carbonic acid (0.125 L, 0.120 M, pKa1 = 6.35, pKa2 = 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 (H*A*) and H_2O, and the dominant equilibrium is H*A* + H_2O A^- + H_3O^+.

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

Reaction: H_2O +	H*A*	A^- +	H_3O^+
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:

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

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

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

(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:

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

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

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

(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^-) and H_2O and the dominant equilibrium is

$$A^{2-} + H_2O \qquad HA^- + OH^-$$

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

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:

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

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

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

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

$[OH^-]_{eq} =$ 4.9×10^{-3} M

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

25. (2 points)

Balance the following redox reaction in acidic conditions:

$$\mathbf{HClO_2 + Co \rightarrow Cl_2 + Co^{2+}}$$

$$\mathbf{BrO_3^- + Fe^{2+} \rightarrow Br^- + Fe^{3+}}$$

$HClO_2 \rightarrow Cl_2$

(1) Multiply $HClO_2$ by 2; (2) add 4 H_2O on the right; (3) add 6 H_3O^+ on the left and 6 H_2O on the right; (4) add 6 e^- on the left:

$2\ HClO_2 + 6\ H_3O^+ + 6\ e^- \rightarrow Cl_2 + 10\ H_2O$

$Co \rightarrow Co^{2+}$

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

Multiply the second reaction by 3 and add:

$$\begin{array}{l} [2\,HClO_2 + 6\,H_3O^+ + 6\,e^- \rightarrow Cl_2 + 10\,H_2O] \\ +\,3\,[Co \rightarrow Co^{2+} + 2\,e^-] \\ \hline 2\,HClO_2 + 3\,Co + 6\,H_3O^+ \rightarrow Cl_2 + 3\,Co^{2+} + 10\,H_2O \end{array}$$

$BrO_3^- \rightarrow Br^-$

(1) Br is balanced; (2) add 3 H_2O on the right; (3) add 6 H_3O^+ on the left and 6 H_2O on the right; (4) add 6 e^- on the left: $BrO_3^- + 6\ H_3O^+ + 6\ e^- \rightarrow Br^- + 9\ H_2O$

$Fe^{2+} \rightarrow Fe^{3+}$

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

$$\begin{array}{l} [BrO_3^- + 6\,H_3O^+ + 6\,e^- \rightarrow Br^- + 9\,H_2O] \\ +\,6\,[Fe^{2+} \rightarrow Fe^{3+} + e^-] \\ \hline BrO_3^- + 6\,Fe^{2+} + 6\,H_3O^+ \rightarrow Br^- + 6\,Fe^{3+} + 9\,H_2O \end{array}$$

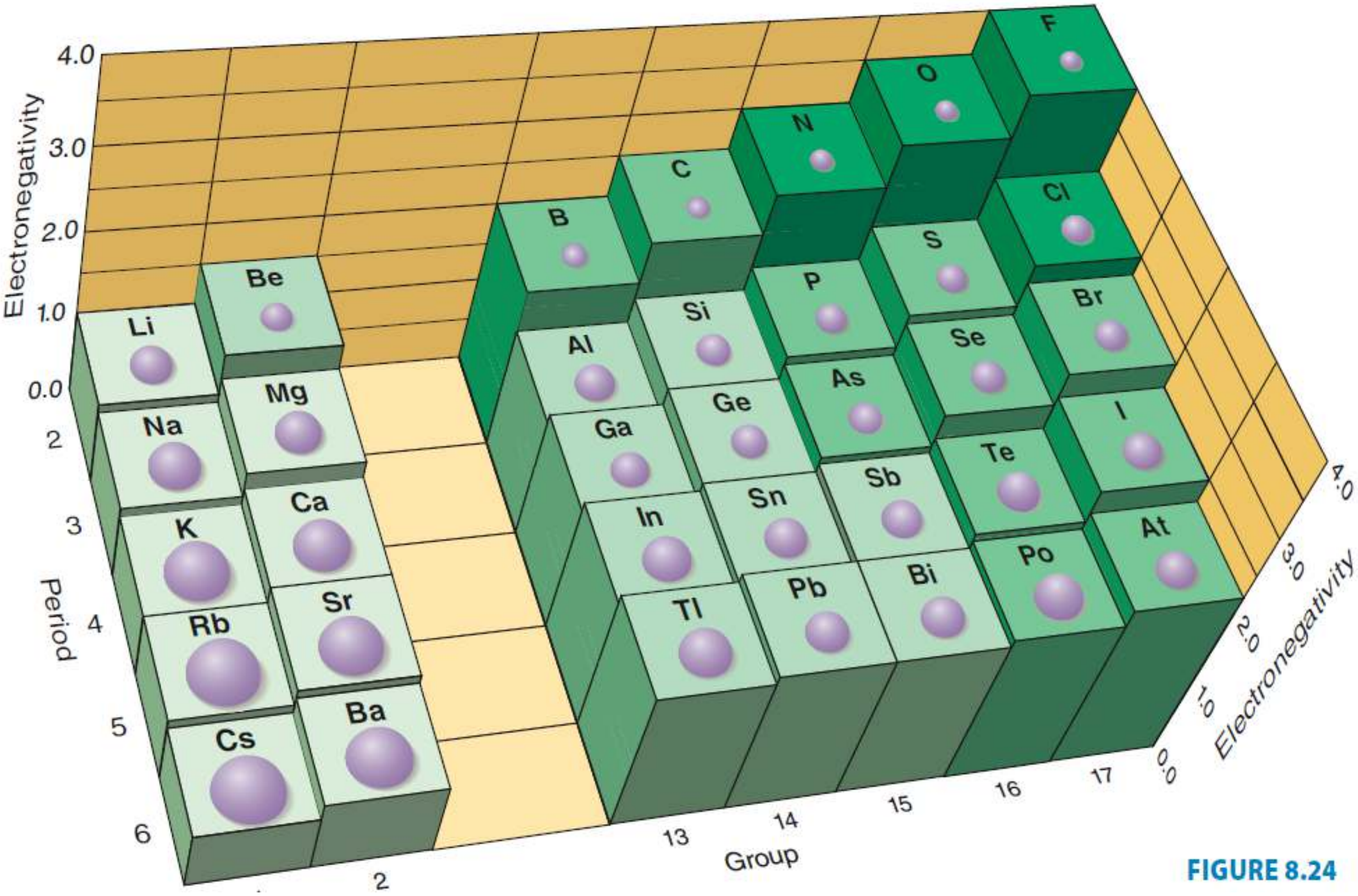
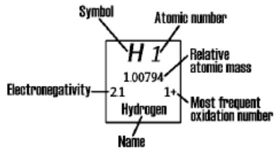

Electronegativity
4.0
3.0
2.0
1.0
0.0
Period
2
3
4
5
6
Li
Be
Na
Mg
K
Ca
Rb
Sr
Cs
Ba
B
C
N
O
F
Al
Si
P
S
Cl
Ga
Ge
As
Se
Br
In
Sn
Sb
Te
I
Tl
Pb
Bi
Po
At
2
13
14
15
16
17
Group
0.0
1.0
2.0
3.0
4.0
Electronegativity

FIGURE 8.24

Mokeur's Periodic table of the elements



	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8	9 VIII	10	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	*H 1* 1.00794 2.1 1+ Hydrogen																	*He 2* 4.002602 - - Helium
2	**Li 3** 6.941 1.0 1+ Lithium	**Be 4** 9.012182 1.5 2+ Beryllium											**B 5** 10.811 2.0 3+ Boron	**C 6** 12.011 2.5 4+,4- Carbon	*N 7* 14.00674 3.0 3+,3- Nitrogen	*O 8* 15.9994 3.5 2- Oxygen	*F 9* 18.9984032 4.0 1- Fluorine	*Ne 10* 20.1797 - - Neon
3	**Na 11** 22.989768 0.9 1+ Sodium	**Mg 12** 24.3050 1.2 2+ Magnesium											**Al 13** 26.981539 1.5 3+ Aluminum	**Si 14** 28.0855 1.8 4+ Silicon	**P 15** 30.973762 2.1 5+ Phosphorus	**S 16** 32.066 1.8 4+ Sulfur	*Cl 17* 35.4527 3.0 1- Chlorine	*Ar 18* 39.948 - - Argon
4	**K 19** 39.0983 0.8 1+ Potassium	**Ca 20** 40.078 1.0 2+ Calcium	**Sc 21** 44.955910 1.3 3+ Scandium	**Ti 22** 47.88 1.5 4+ Titanium	**V 23** 50.9415 1.6 5+ Vanadium	**Cr 24** 51.9961 1.6 3+ Chromium	**Mn 25** 54.93805 1.5 2+ Manganese	**Fe 26** 55.847 1.8 3+ Iron	**Co 27** 58.9332 1.8 2+ Cobalt	**Ni 28** 58.6934 1.8 2+ Nickel	**Cu 29** 63.546 1.9 2+ Copper	**Zn 30** 65.39 1.6 2+ Zinc	***Ga 31*** 69.723 1.6 3+ Gallium	**Ge 32** 72.61 1.6 4+ Germanium	**As 33** 74.92159 2.0 3+,3- Arsenic	**Se 34** 78.96 2.4 4+ Selenium	***Br 35*** 79.904 2.8 1- Bromine	*Kr 36* 83.80 - - Krypton
5	**Rb 37** 85.4678 0.8 1+ Rubidium	**Sr 38** 87.62 1.0 2+ Strontium	**Y 39** 88.90585 1.3 3+ Yttrium	**Zr 40** 91.224 1.4 4+ Zirconium	**Nb 41** 92.90638 1.6 5+ Niobium	**Mo 42** 95.94 1.8 6+ Molybdenum	Tc 43 98.9063 1.9 7+ Technetium	**Ru 44** 101.57 2.2 3+,4+ Ruthenium	**Rh 45** 102.9055 2.2 3+ Rhodium	**Pd 46** 106.42 2.2 2+ Palladium	**Ag 47** 107.8682 1.9 1+ Silver	**Cd 48** 112.411 1.7 2+ Cadmium	**In 49** 114.82 1.7 3+ Indium	**Sn 50** 118.71 1.8 4+ Tin	**Sb 51** 121.757 1.9 3+,3- Antimony	**Te 52** 127.60 2.1 4+ Tellurium	**I 53** 126.90447 2.5 1- Iodine	*Xe 54* 131.29 - - Xenon
6	***Cs 55*** 132.90543 0.7 1+ Cesium	**Ba 56** 137.327 0.9 2+ Barium	**La 57** 138.9055 1.1 3+ Lanthanum	**Hf 72** 178.49 1.3 4+ Hafnium	**Ta 73** 180.9479 1.5 5+ Tantalum	**W 74** 183.85 1.7 6+ Tungsten	**Re 75** 186.207 1.9 7+ Rhenium	**Os 76** 190.2 2.2 4+ Osmium	**Ir 77** 192.22 2.2 4+ Iridium	**Pt 78** 195.08 2.2 4+ Platinum	**Au 79** 196.96654 2.4 3+ Gold	***Hg 80*** 200.59 1.9 2+ Mercury	**Tl 81** 204.3833 1.8 1+ Thallium	**Pb 82** 207.2 1.8 2+ Lead	**Bi 83** 208.98037 1.9 3+ Bismuth	**Po 84** 208.9824 2.0 2+ Polonium	**At 85** 209.9871 2.2 1- Astatine	*Rn 86* 222.0176 - - Radon
7	***Fr 87*** 223.0197 0.7 1+ Francium	**Ra 88** 226.0254 0.9 2+ Radium	**Ac 89** 227.0278 1.1 3+ Actinium	Rf 104 261.11 - - Rutherfordium	Db 105 262.11 - - Dubnium	Sg 106 263.12 - - Seaborgium	Bh 107 262.12 - - Bohrium	Hs 108 264 - - Hassium	Mt 109 266.1378 - - Meitnerium	Uun 110 269 - - Ununnilium	Uuu 111 272 - - Unununium	Uu 112 277 - - Ununbium		Uuq 114 289 - - Ununquadium		Uuh 116 289 - - Ununhexium		Uuo 118 293 - - Ununoctium

6	**Ce 58** 140.115 1.1 3+ Cerium	**Pr 59** 140.90765 1.1 3+ Praseodymium	**Nd 60** 144.24 1.1 3+ Neodymium	Pm 61 144.9127 1.1 3+ Promethium	**Sm 62** 150.36 1.2 3+ Samarium	**Eu 63** 151.965 1.2 3+ Europium	**Gd 64** 157.25 1.2 3+ Gadolinium	**Tb 65** 168.92534 1.2 3+ Terbium	**Dy 66** 162.50 1.2 3+ Dysposium	**Ho 67** 164.93032 1.2 3+ Holmium	**Er 68** 167.26 1.2 3+ Erbium	**Tm 69** 168.93421 1.2 3+ Thulium	**Yb 70** 173.04 1.1 3+ Ytterbium	**Lu 71** 174.967 1.2 3+ Lutetium
7	**Th 90** 232.0381 1.3 4+ Thorium	**Pa 91** 231.03588 1.5 5+ Protactinium	**U 92** 238.0289 1.4 6+ Uranium	Np 93 237.082 1.3 5+ Neptunium	Pu 94 244.0642 1.3 4+ Plutonium	Am 95 243.0614 1.3 3+ Americium	Cm 96 247 1.3 3+ Curium	Bk 97 247.0703 1.3 3+ Berkelium	Cf 98 251.0796 1.3 3+ Californium	Es 99 252.03 1.3 - Einsteinium	Fm 100 257.0951 1.3 - Fermium	Md 101 258.01 1.3 - Mendelevium	No 102 259.1009 1.3 - Nobelium	Lr 103 260.1053 - - 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.

Data For Water

Density = 1.00 g/mL (at 25°C)
s = 2.13 J g^{-1} K^{-1} (solid)
s = 4.184 J g^{-1} K^{-1} (liquid) $\Delta H°_{fus}$ = 6.02 kJ mol^{-1}
s = 2.01 J g^{-1} K^{-1} (gas) $\Delta H°_{vap}$ = 40.7 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^3 = 1 mL 1000 mL = 1 L 1000 L = 1 m^3

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

Gas Laws

$PV = nRT$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

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

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

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

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

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

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

Equilibrium

$K_P = K_C(RT^{)\Delta n}$

Acid/Base

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

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

$pH + pOH = 14$

$K_a \times K_b = K_W$

$pH = pK_a + \log [A^-]/[HA]$

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

Thermochemistry

$\Delta U = q + W$

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

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

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

$q = ms\Delta T$

$\Delta H_{rxn}° = \sum n\Delta H_f°(pdts) - \sum n\Delta H_f°(rxts)$

The atom

$E = h\nu$

$c = \nu\lambda$

$E = -B/n^2$

Kinetics

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

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

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

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

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