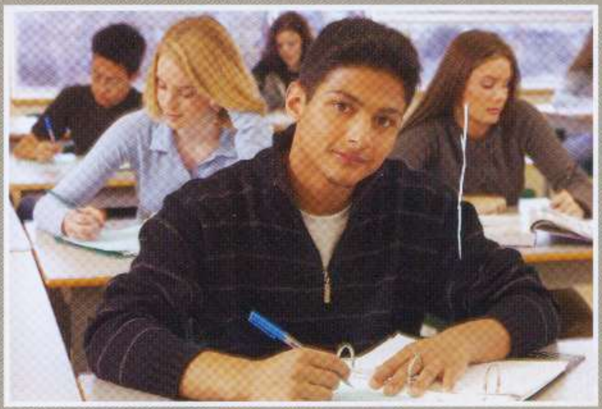
Chemistry 12

UNIVERSITY PREPARATION (SCH4U)

THE KEY

STUDY GUIDE



- 100% aligned with the Ontario curriculum
- Includes unit reviews, practice questions, and tests
- Contains answers, explanations, and detailed solutions
- Complements classroom instruction
- Reviewed by respected Ontario educators

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Ontario
Edition

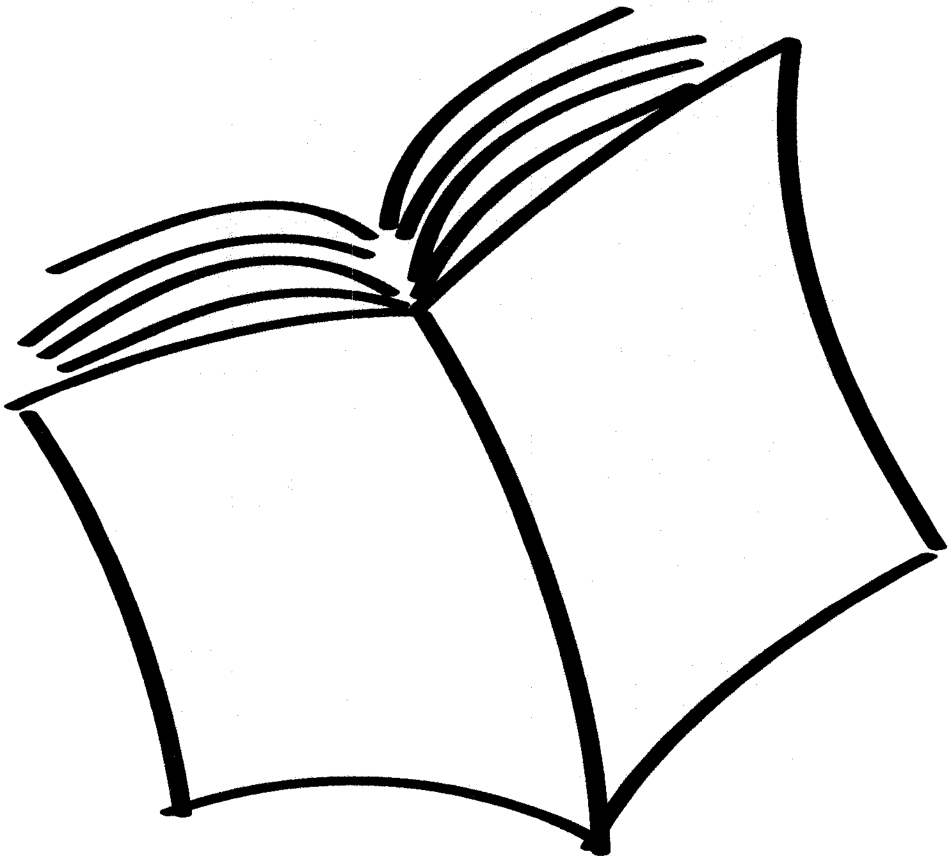
THE KEY

STUDENT STUDY GUIDE

Chemistry 12

University Preparation (SCH4U)

THE KEY student study guide is designed to help students achieve success in school. The content in each study guide is 100% curriculum aligned and serves as an excellent source of material for review and practice. To create this book, teachers, curriculum specialists, and assessment experts have worked closely to develop the instructional pieces that explain each of the key concepts for the course. The practice questions and sample tests have detailed solutions that show problem-solving methods, highlight concepts that are likely to be tested, and point out potential sources of errors. ***THE KEY*** is a complete guide to be used by students throughout the school year for reviewing and understanding course content, and to prepare for assessments.



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Dedicated to the memory of Dr. V. S. Rao

THE KEY—Chemistry 12 University Preparation

THE KEY consists of the following sections:

KEY Tips for Being Successful at School gives examples of study and review strategies. It includes information about learning styles, study schedules, and note taking for test preparation.

Class Focus includes a unit on each area of the curriculum. Units are divided into sections, each focusing on one of the specific expectations, or main ideas, that students must learn about in that unit. Examples, definitions, and visuals help to explain each main idea. Practice questions on the main ideas are also included. At the end of each unit is a test on the important ideas covered. The practice questions and unit tests help students identify areas they know and those they need to study more. They can also be used as preparation for tests and quizzes. Most questions are of average difficulty, though some are easy and some are hard—the harder questions are called *Challenger Questions*. Each unit is prefaced by a ***Table of Correlations***, which correlates questions in the unit (and in the practice tests at the end of the book) to the specific curriculum expectations. Answers and solutions are found at the end of each unit.

KEY Strategies for Success on Tests helps students get ready for tests. It shows students different types of questions they might see, word clues to look for when reading them, and hints for answering them.

Practice Tests includes one to two tests based on the entire course. They are very similar to the format and level of difficulty that students may encounter on final tests. In some regions, these tests may be reprinted versions of official tests, or reflect the same difficulty levels and formats as official versions. This gives students the chance to practice using real-world examples. Answers and complete solutions are provided at the end of the section.

For the complete curriculum document (including specific expectations along with examples and sample problems), http://www.edu.gov.on.ca/eng/curriculum/secondary/science1112curr.pdf.

THE KEY *Study Guides* are available for many courses. Check www.castlerockresearch.com for a complete listing of books available for your area.

For information about any of our resources or services, please call Castle Rock Research at 905.625.3332 or visit our website at http://www.castlerockresearch.com.

At Castle Rock Research, we strive to produce an error-free resource. If you should find an error, please contact us so that future editions can be corrected.

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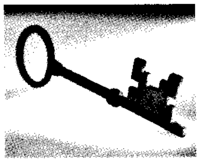
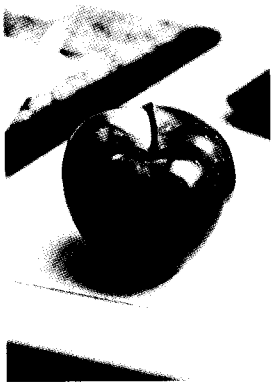
Success at School



KEY FACTORS CONTRIBUTING TO SCHOOL SUCCESS

In addition to learning the contents of your courses, there are some other things that you can do to help you do your best at school. Some of these strategies are listed below.

- **ATTEND SCHOOL REGULARLY** so you do not miss any classes, notes, or important activities that will help you learn.
- **KEEP A POSITIVE ATTITUDE.** Always reflect on what you can already do and what you already know.
- **BE PREPARED TO LEARN.** Have the necessary materials (pencils, pens, notebooks, and other required materials) with you in class.
- **COMPLETE ALL OF YOUR ASSIGNMENTS.** Do your best to finish all of your assignments. Even if you know the material well, practice will reinforce your knowledge. If an assignment or question is difficult for you, work through it as far as you can so your teacher can see exactly where you are having difficulty.
- **SET SMALL GOALS** for yourself when you are learning new material. For example, when learning formulas, do not try to learn everything in one night. Work on only one formula each study session. When you understand one particular formula and have memorized it, move on to another one. Continue this process until you have learned and memorized all of the required formulas.
- **REVIEW YOUR CLASSROOM WORK** regularly at home to be sure you understand the material you learned in class.
- **ASK YOUR TEACHER FOR HELP** when you do not understand something or when you are having difficulty completing your assignments.
- **GET PLENTY OF REST AND EXERCISE.** Concentrating in class is hard work. It is important to be well-rested and have time to relax and socialize with your friends. This helps you to keep a positive attitude about your school work.
- **EAT HEALTHY MEALS.** A balanced diet keeps you healthy and gives you the energy you need for studying at school and at home.



HOW TO FIND YOUR LEARNING STYLE

Every student has a certain manner in which it seems easier for him or her to learn. The manner in which you learn best is called your learning style. By knowing your learning style, you can increase your success at school. Most students use a combination of learning styles.

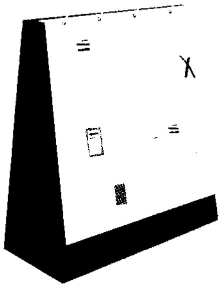
Do you know what type of learner you are? Read the following descriptions. Which of these common learning styles do you use most often?

- **Do you need to say things out loud?** You may learn best by saying, hearing, and seeing words. You are probably really good at memorizing dates, places, names, and facts. To learn the steps in a process, a formula, or the actions that lead up to a significant event, you may need **to write them and then read them out loud**.

- **Do you need to read or see things?** You may learn best by looking at and working with pictures. You are probably really good at puzzles, imagining things, and reading maps and charts. You may need to use strategies like **mind mapping and webbing** to organize your information and study notes.

- **Do you need to draw or write things down?** You may learn best by touching, moving, and figuring things out using manipulatives. You are probably really good at physical activities and learning through movement. You may need to **draw your finger over a diagram** to remember it, ***tap out* the steps** needed to solve a problem, or ***feel* yourself writing** or typing a formula.

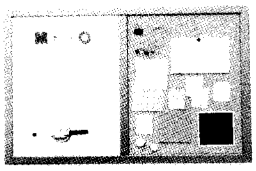
SCHEDULING STUDY TIME



You should review your class notes regularly to be sure you have a clear understanding of all the new material you learned. Reviewing your lessons on a regular basis helps you to learn and remember ideas and concepts. It also reduces the quantity of material you need to study prior to a test. Creating a study schedule will help you to make the best use of your time.

Regardless of the type of study schedule you use, you may want to consider the following strategies for making the most of your study time and effort:

- Organize your work so you begin with the most challenging material first.
- Divide the subject content into small, manageable chunks.
- Alternate regularly between your different subjects and types of study activities in order to maintain your interest and motivation.
- Make a daily list with the headings *must do*, *should do*, and *could do*.
- Begin each study session by quickly reviewing what you studied the day before.
- Maintain your usual routine of eating, sleeping, and exercising to help you concentrate better for extended periods of time.

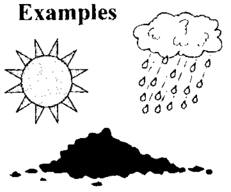
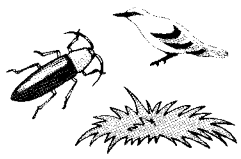


CREATING STUDY NOTES

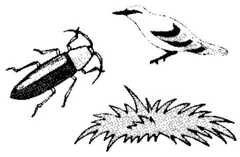
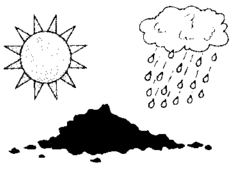
MIND-MAPPING OR WEBBING

- Use the key words, ideas, or concepts from your class notes to create a *mind map* or *mind web*, which is a diagram or visual representation of the given information. A mind map or web is sometimes referred to as a *knowledge map*.
- Write the key word, concept, theory, or formula in the centre of your page.
- Write down related facts, ideas, events, and information and then link them to the central concept.
- The following examples of a Frayer Model illustrate how this technique can be used to study scientific vocabulary.

Abiotic

Definition – Non-living components of an ecosystem	**Notes** – The abiotic factors of an area determine biotic factors that can live there
Examples 	**Non-examples** 

Biotic

Definition – Living components of an ecosystem	**Notes** – The types of living things that can live in an area depends on the ability of the abiotic factors to meet their needs
Examples 	**Non-examples** 

INDEX CARDS

To use index cards while studying, follow these steps:

- Write a key word or question on one side of an index card.
- On the other side, write the definition of the word, answer to the question, or any other important information you want to remember.

What is the difference between heat and thermal energy?

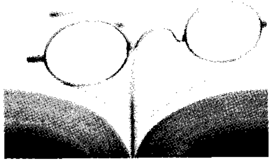
What is the difference between heat and thermal energy?

Thermal energy is the total energy of the particles in a solid, liquid, or gas. Heat is the amount of the thermal energy that is transferred between two objects.

SYMBOLS AND STICKY NOTES—IDENTIFYING IMPORTANT INFORMATION

- Use symbols to mark your class notes. For example, an exclamation mark (!) might be used to point out something that must be learned well because it is a very important idea. A question mark (?) may highlight something you are not certain about, and a diamond (◊) or asterisk (*) could mark interesting information you want to remember.
- Use sticky notes to mark a page in a book that contains an important diagram, formula, or explanation.

KEY STRATEGIES FOR REVIEWING



Reviewing textbook material, class notes, and handouts should be an ongoing activity. Spending time reviewing becomes more critical when you are preparing for tests. You may find some of the following review strategies useful when studying during your scheduled study time.

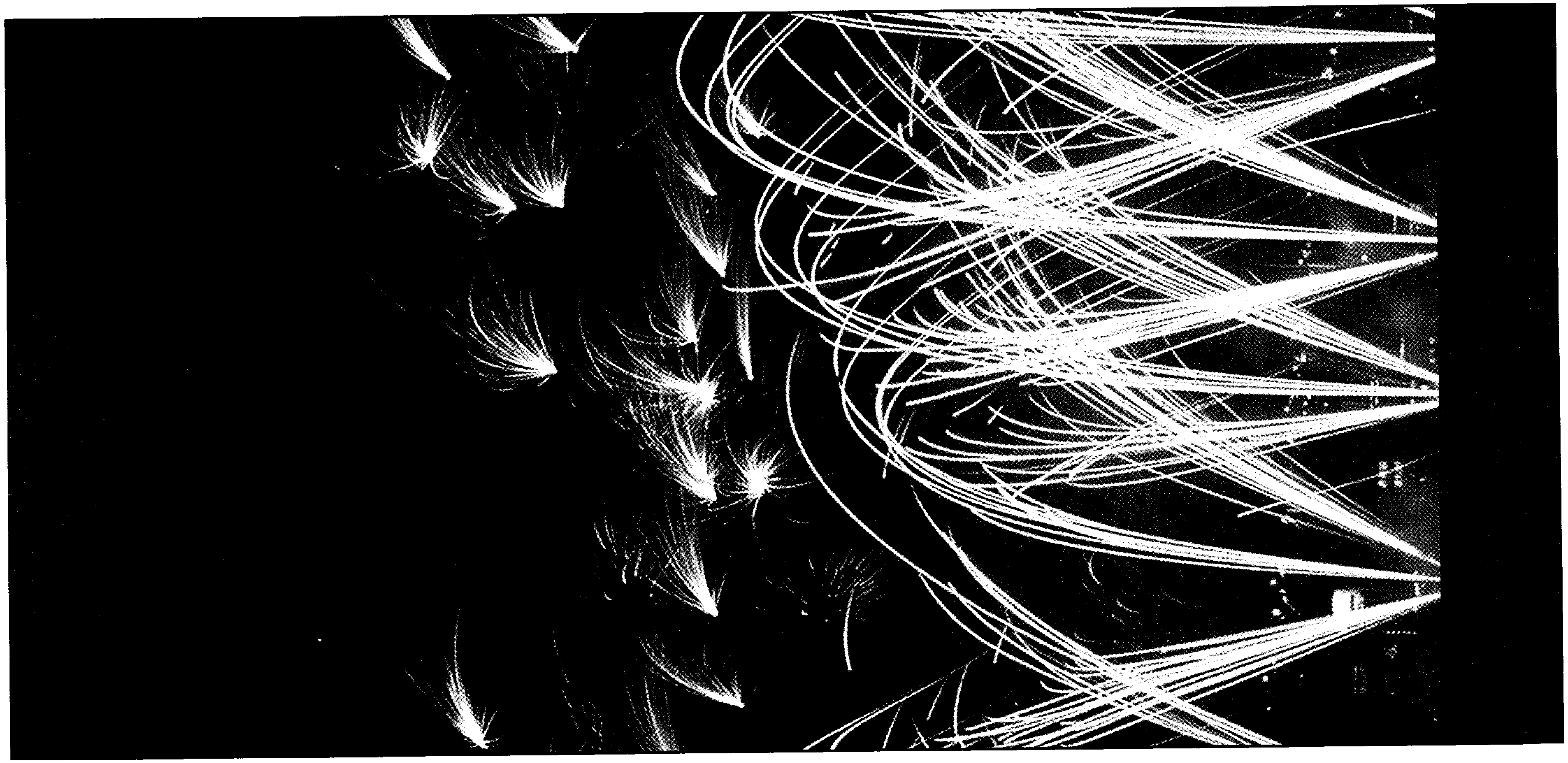
- Before reviewing a unit, note the headings, charts, graphs, and chapter questions.
- Highlight mathematical key concepts, vocabulary, definitions, and formulas.
- Carefully read over each step in a procedure.
- Draw a picture or diagram to help make the concept clearer.

KEY STRATEGIES FOR SUCCESS—A CHECKLIST

Review, review, review: that is a huge part of doing well at school and preparing for tests. Below is a checklist for you to keep track of how many suggested strategies for success you use. Read each question and then put a check mark (✓) in the correct column. Look at the questions for which you have checked the *No* column. Think about how you might try using some of these strategies to help you do your best at school.

***KEY* Strategies for Success**	**Yes**	**No**
Do you attend school regularly?		
Do you know your personal learning style—how you learn best?		
Do you spend 15 to 30 minutes each day reviewing your notes?		
Do you study in a quiet place at home?		
Do you clearly mark the most important ideas in your study notes?		
Do you use sticky notes to mark texts and research books?		
Do you practice answering multiple-choice and written-response questions?		
Do you ask your teacher for help when you need it?		
Do you maintain a healthy diet and sleep routine?		
Do you participate in regular physical activity?		

Energy Changes



Energy Changes and Rates of Reaction

Table of Correlations

Specific Expectation	Practice Questions	Self Test Questions
12.2.1 Understanding Basic Concepts		
12.2.1.1 *compare the energy changes resulting from physical changes, chemical reactions, and nuclear reactions (fission and fusion)*	1	1
12.2.1.2 *explain Hess's law, using examples*	2, 3	2
12.2.1.3 *describe, with the aid of a graph, the rate of reaction as a function of the change of concentration of a reactant or product with respect to time; express the rate of reaction as a rate law equation (first- or second-order reactions only); and explain the concept of half-life for a reaction*	4, 5a, 5b	3
12.2.1.4 *analyse simple potential energy diagrams of chemical reactions*	6, 7	4
12.2.1.5 *demonstrate understanding that most reactions occur as a series of elementary steps in a reaction mechanism*	8	5
12.2.2 Developing Skills of Inquiry and Communication		
12.2.2.1 *use appropriate scientific vocabulary to communicate ideas related to the energetics of chemical reactions*	9	
12.2.2.2 *write thermochemical equations, expressing the energy change as a ΔH value or as a heat term in the equation*	10, 11	6
12.2.2.3 *determine heat of reaction using a calorimeter, and use the data obtained to calculate the enthalpy change for a reaction*	12, 13, 14, 15	7, 8
12.2.2.4 *apply Hess's law to solve problems, including problems that involve data obtained through experimentation*	16, 17, 18, 19	16, 17
12.2.2.5 *calculate heat of reaction using tabulated enthalpies of formation*	20, 21, 22	9, 10
12.2.2.6 *determine through experimentation a rate of reaction, and measure the effect on it of temperature, concentration, and catalysis*	23, 24a, 24b	11
12.2.3 Relating Science to Technology, Society, and the Environment		
12.2.3.1 *compare conventional and alternative sources of energy with respect to efficiency and environmental impact*	25, 26	12
12.2.3.2 *describe examples of technologies that depend on exothermic or endothermic changes*	27	13
12.2.3.3 *describe the use of catalysts in industry and in biochemical systems on the basis of information gathered from print and electronic sources*	28	14
12.2.3.4 *describe examples of slow chemical reactions, rapid reactions, and reactions whose rates can be controlled*	29	15

12.2.1.1 *compare the energy changes resulting from physical changes, chemical reactions, and nuclear reactions (fission and fusion)*

RELATIVE ENTHALPIES

The term enthalpy describes the total energy of a system at constant pressure. It includes various types of energy such as heat, kinetic, and potential energy. Since only heat energy is measurable, energy is often measured as a change in heat energy. This is called enthalpy change or ΔH. If a reaction gives off more heat energy than it absorbs, it is an exothermic reaction. If the reaction absorbs more energy that it releases, it is termed an endothermic reaction. The amount of enthalpy change or heat energy a reaction will produce or absorb is dependent on the type of reaction.

Physical changes produce the least enthalpy change as only intermolecular forces of attraction are broken or formed. Chemical changes involve the breaking and forming of chemical bonds between atoms. In these reactions, energy is needed to break reactant intramolecular bonds. Energy is then released when new bonds are formed. The reason more enthalpy change is observed in chemical changes is that the attraction between atoms in a chemical bond is greater than the attraction between molecules. Finally, nuclear reactions yield the greatest enthalpy change. Internuclear forces are broken in the nuclear fission reaction, or two nuclei are forced to merge in the nuclear fusion reaction. The forces of attraction within the nucleus are even greater than those in chemical bonds, thereby yielding the highest enthalpy change.

The relative enthalpy change of these three types of reactions can be summarized as follows:

PHYSICAL < CHEMICAL ≪ NUCLEAR

Practice

1. When energy changes are arranged in **decreasing** order of magnitude, the order is
   A. phase, chemical, nuclear fusion
   B. nuclear fusion, phase, chemical
   C. chemical, nuclear fusion, phase
   D. nuclear fusion, chemical, phase

12.2.1.2 *explain Hess's law, using examples*

HESS'S LAW

The enthalpy change of a net reaction is the sum of the enthalpies of the reactions combined to make the net reaction.

Consider:

(1) $A + B \rightarrow C + D \quad \Delta H = \Delta H_1$

(2) $C + D \rightarrow E + F \quad \Delta H = \Delta H_2$

Then, the enthalpy change of the reaction

(net) $A + B \rightarrow E + F$
$\Delta H_{net} = \Delta H_1 + \Delta H_2$

Alternatively, Hess's law exemplifies the principle that, regardless of the route from reactants to products, the enthalpy change of a process is the difference in enthalpy between the products and reactants.

Example

Find the ΔH for $Z_2Y_3 + 3XY \rightarrow 3XY_2 + 2Z$ using the equations with their given enthalpy changes.

Strategy: Look at the desired equation and rearrange the given equations to produce the compounds with the correct coefficients. Do not worry about pure elements—they will work out automatically. Do worry about extra compounds that have been added—they need to be cancelled out.

A. $2XY + Y_2 \rightarrow 2XY_2 \Delta H = -525.0$ kJ

B. $4Z + 3Y_2 \rightarrow 2Z_2Y_3 \Delta H = -742.2$ kJ

Multiply equation A by $\frac{3}{2}$ to change the coefficient of XY_2 to 3 $\Rightarrow 3XY + \frac{3}{2}Y_2 \rightarrow 3XY_2$

$\Delta H = -787.5$ kJ

Flip equation B and divide by 2 to get 1 Z_2Y_3 on the left $\Rightarrow Z_2Y_3 \rightarrow 2Z + \frac{3}{2}Y_2 \Delta H = +371.1$ kJ (note sign change)

Add the last two equations and you will see that they add to the desired equation, and that the $\Delta H = -787.5\text{ kJ} + 371.1\text{ kJ} = 416.4\text{ kJ}$

Practice

Use the following information to answer the next question.

I. $A + B \rightarrow C + D$

II. $A + B \rightarrow R$
$R \rightarrow C + D$

2. Products C and D can be formed from reactants A and B through either a single step process (I) or a two-step process (II). Which of the following statements about the enthalpy changes of the two processes is **true**?

A. The enthalpy changes for both processes are the same.

B. The enthalpy change for process I is greater than that of process II.

C. The enthalpy change for process II is greater than that of process I.

D. There is not enough information to determine the relationship between process I and process II.

Open Response

Use the following information to answer the next question.

The reaction of ethene with hydrogen gas produces ethane as follows:

$$C_2H_{4(g)} + H_{2(g)} \rightarrow C_2H_{6(g)}$$

3. Can the enthalpy of the given reaction be determined using the following data? Explain why or why not using Hess's law.

Intermediate Reactions:

$$C_2H_{4(g)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 2H_2O_{(l)}$$
$$\Delta H = -1\,411\text{ kJ}$$

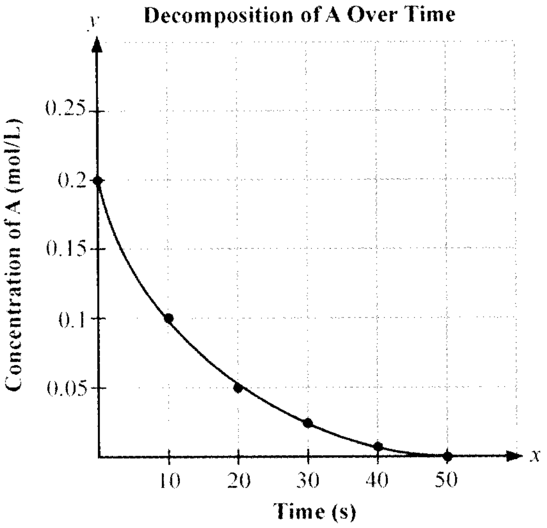
$$C_2H_{6(g)} + \frac{7}{2}O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(l)}$$
$$\Delta H = \quad 1\,560\text{ kJ}$$

$$H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(g)}\Delta H = -241.8\text{ kJ}$$

12.2.1.3 *describe, with the aid of a graph, the rate of reaction as a function of the change of concentration of a reactant or product with respect to time; express the rate of reaction as a rate law equation (first- or second-order reactions only); and explain the concept of half-life for a reaction*

RATE OF REACTION AND RATE LAW

The rate of reaction is the change in concentration of the reactants or products over time. The given graph shows a general representation of the change in concentration over time for a chemical reaction.



The rate of this reaction at any time t, can be found by taking the tangent to the curve, and then finding the slope of the tangent.

Finding the rate: $r_{\text{av}} = \dfrac{\Delta C}{\Delta t}$

So mathematically, the rate of a reaction is proportional to the concentration of the reactants. $r \propto [X][Y]$ or $r = k[X][Y]$

Knowing this, a rate law can be determined giving the rate of a reaction without needing to derive it from a graph. The example of the reaction of hydrogen and iodine is given.

$H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)}$

Trial	$[H_2]$	$[I_2]$	Rate
1	0.1	0.1	1×10^{-2}
2	0.1	0.2	2×10^{-2}
3	0.2	0.2	8×10^{-2}

From the table, it is observed that:

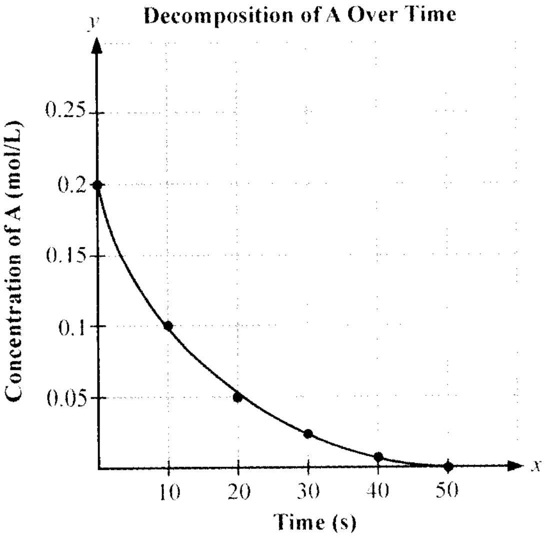
1. When $[H_2]$ is kept constant (trials 1 and 2) and $[I_2]$ is doubled, then the rate is also doubled. Therefore, $r = k[I_2]$.
2. When $[I_2]$ is kept constant (trials 2 and 3) and $[H_2]$ is doubled, then the rate is increased by a factor of 4. Therefore, $r = k[H_2]^2$.

Thus, the rate law for the reaction of iodine and hydrogen gas is: $r = k[I_2][H_2]^2$. The rate order is the sum of the exponents in the rate law, which in this case would be 3. In summary:

Rate $= k[X]^m[Y]^n$

Order $= m + n$

Finally, the half-life is the amount of time needed for a reactant mass or concentration to decrease by $\frac{1}{2}$ its initial value. In the decomposition of compound A, the half-life can be determined from the graph of concentration vs. time.

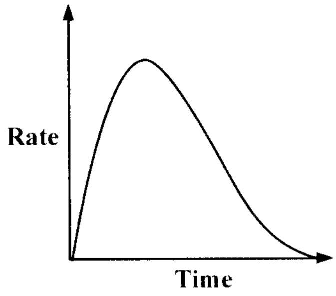


Since the product of the rate constant and half-life is constant for any **first-order** reaction, the half-life of a **first-order** reaction can be calculated using: $t_{\frac{1}{2}} = \frac{0.693}{k}$.

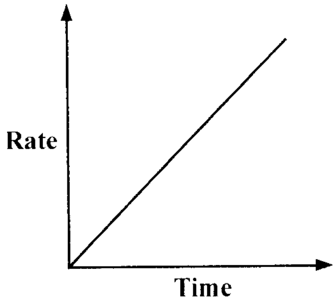
Practice

4. Which of the following graphs represents the typical mathematical relationship between reaction rate and time?

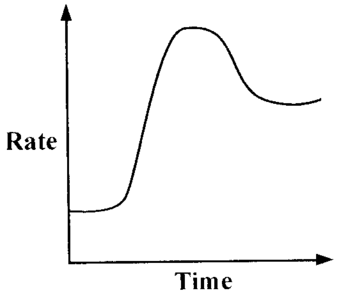
A.



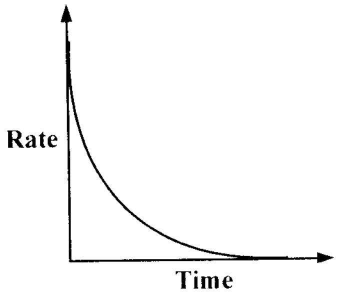
B.



C.



D.



Use the following information to answer the next multipart question.

5. The hydrolysis of butyl chloride takes place as

$C_4H_9Cl + H_2O \rightarrow C_4H_9OH + HCl$

The concentration of C_4H_9Cl at different points of time during the reaction is given in the table.

t (s)	$[C_4H_9Cl]$(mol L^{-1})
0	0.1000
50	0.0905
100	0.0820
150	0.0741
200	0.0671
300	0.0549
400	0.0439
700	0.0210
800	0.0170

Part A

Open Response

What is the rate of the reaction for the time period 300 s to 400 s?

Part B

Open Response

What is the instantaneous rate of the reaction at $t = 600$ s?

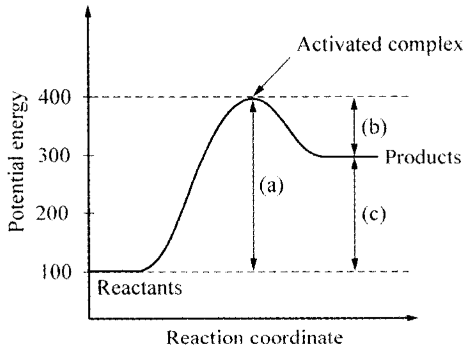
12.2.1.4 *analyse simple potential energy diagrams of chemical reactions*

ENERGY DIAGRAMS

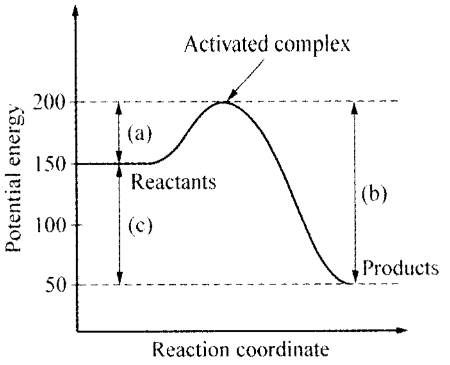
The following two diagrams illustrate the potential energy of reactants, products, and the activated complex in a chemical reaction. In both diagrams,

a. represents the forward activation energy,
b. represents the reverse activation energy, and
c. represents the enthalpy change, ΔH.

Endothermic reaction

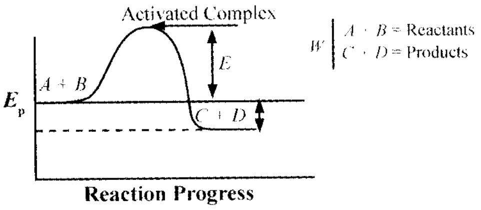


Exothermic reaction



Practice

Use the following information to answer the next question.



Energy diagram of an exothermic reaction

6. The E variable in the energy diagram represents

A. activation energy

B. potential energy

C. thermal energy

D. kinetic energy

Open Response

Use the following information to answer the next question.

The combustion of sugar in a bomb calorimeter is similar to the oxidation of sugar in the body. A student ate three sugar cubes with masses of 6.84 g, 6.75 g, and 6.79 g, respectively.

7. Draw and label a potential energy diagram representing the molar enthalpy of oxidation of sugar in the body.

12.2.1.5 *demonstrate understanding that most reactions occur as a series of elementary steps in a reaction mechanism*

REACTION MECHANISM

Most reactions are a combination of several elementary steps that when combined, add up to the overall reaction. For example:

The reaction: $2NO_2 + F_2 \rightarrow 2NO_2F$

Follows the mechanism:

i) $NO_2 + F_2 \rightarrow NO_2F + F$ (slow)

ii) $NO_2 + F \rightarrow NO_2F$ (fast)

The rate law of a reaction is consistent with the rate-determining step and the exponents in the rate law are the same as the coefficients in the rate-determining step. So, for the given reaction, the rate law is:

$r = k[NO_2][F_2]$

If a reaction occurs in a single step, it is referred to as an elementary process. Only the decomposition of a compound and binary collisions are elementary processes.

Practice

Use the following information to answer the next question.

Formation of hydrogen bromide takes place as
$H_2 + Br_2 \rightleftharpoons 2HBr$

The rate law for the reaction is

$$\text{Rate} = k_2\left(\frac{k_1}{k_{-1}}\right)^{\frac{1}{2}}[H_2][Br_2]^{\frac{1}{2}}$$

8. What can be concluded from the rate law of reaction?

A. The given reaction is an elementary reaction.

B. The rate law has been deduced from the stoichiometry of the equation.

C. The given reaction involves more than one-step.

D. The rate law for the reaction is dependent on the concentration of HBr.

12.2.2.1 *use appropriate scientific vocabulary to communicate ideas related to the energetics of chemical reactions*

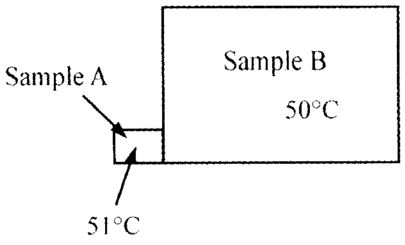
THERMOCHEMICAL VOCABULARY

In order for scientists to communicate effectively with each other, they must use the same terminology. Each term in their vocabulary must have the same, precise meaning for all of the scientific community. There are many terms associated with the study of energy in chemical systems.

The study of energy changes and transformations is called **thermodynamics**. If that study is associated with chemical reactions, it is called **thermochemistry**.

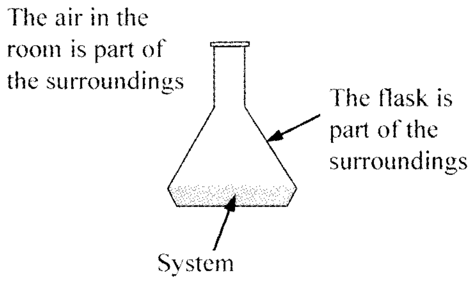
The particles of a sample of any substance are in constant motion. The energy of this motion is called **thermal energy**. This energy can be transferred from one sample to another in the form of **heat** (q). In order for heat to be transferred, it must flow into another substance or substances. If heat flows out of a substance, the process is called **exothermic**. If it flows into a substance, the process is called **endothermic**. If a substance undergoes an exothermic change, its thermal energy decreases and if it undergoes an endothermic change, its thermal energy increases.

Not all particles in a sample have the same kinetic energy (energy of motion). A measure of the average kinetic energy in a sample is the **temperature** (T) of the sample. If two samples are in contact with each other and if they are at different temperatures, heat will flow from the substance that has the higher temperature, regardless of the thermal energy of the substances.



Sample B, as a result of its greater mass, has more thermal energy than sample A. However, the reason heat will flow from sample A to sample B is that sample A has a greater temperature.

A **system** is any part of the universe that we wish to study. For a chemist, a system is normally a chemical reaction. Everything else is called the **surroundings**.



The law of conservation of energy states that energy can neither be created nor destroyed. This means that when energy is lost by a system, it is gained by the surroundings and when energy is gained by the system, it is lost by the surroundings.

A system that allows matter to be exchanged with the surroundings is called an **open system**. A system that does not allow matter to be exchanged with the surroundings but does allow energy to be exchanged is called a **closed system**.

The term used for the measurement of energy changes in a system is called **calorimetry** and the vessel in which thermochemical reactions are carried out is called a **calorimeter**.

If the same mass of two different substances are given the same amount of heat energy, the temperature of the two substances will not increase by the same amount. The resulting change in temperature depends on the **specific heat capacity** (c) of each substance. Specific heat capacity is defined as the amount of heat required to raise the temperature of one gram of a substance by one degree. If the temperature of a sample were raised by 10°C and the amount of heat added were measured, then the specific heat capacity could be calculated as follows:

$$c = \frac{q}{m\Delta T}$$

where m = mass and ΔT = change in temperature

A more useful form of this equation is the rearranged form, $q = mc\Delta T$.

Example

Calculate the amount of heat required to raise the temperature of 127.4 g of Mg from 22.0°C to 78.0°C.

Solution

Given: $m_{Mg} = 127.4\ \text{g}$

$T_i = 22.0°\text{C}$

$T_f = 78.0°\text{C}$

Required: $q = ?$

We can use the equation, $q = mc\Delta T$. However, we are not given c for magnesium. We must look up the value of c in tables of specific heat capacities. From the tables: $c_{\text{Mg}} = 1.017\ \text{J} \cdot \text{g}^{-1}°\text{C}^{-1}$.

$\Delta T = T_f - T_i$ by definition

$$\begin{aligned} q &= mc\Delta T \\ &= (127.4\ \text{g})(1.017\ \text{J} \cdot \text{g}^{-1}°\text{C}^{-1})(78.0°\text{C} - 22.0°\text{C}) \\ &= 7.26 \times 10^3\ \text{J} \end{aligned}$$

The amount of heat required is 7.26×10^3 J.

When dealing with a chemical reaction, chemists use the concept of **enthalpy** (H) rather than thermal energy. Enthalpy is defined as thermal energy measured at constant pressure. Since reactions are normally carried out in vessels that are open to the atmosphere, the pressure of the reaction mixture is constant. Also, chemists are concerned with changes in enthalpy (ΔH) and not with absolute enthalpy. The enthalpy change for a chemical reaction is denoted by the symbol ΔH_r. The value of ΔH depends on the pressure and temperature of the reaction, so values for ΔH are usually quoted for SATP conditions. To indicate that the value is for SATP conditions, the symbol ΔH_r° is used. This is called the **standard enthalpy of reaction**.

When a chemist includes the value for ΔH in an equation, the equation is referred to as a **thermochemical equation**. There are two ways of writing thermochemical equations. For an exothermic reaction, the heat produced can be considered to be a product of the reaction and can be written on the product side of the equation.

Example

$H_{2(g)} + O_{2(g)} \rightarrow H_2O_{(l)} + 285.8\ \text{kJ}$

Since enthalpy has left the system, and $\Delta H_r^{\circ} = \Delta H_f^{\circ} - \Delta H_i^{\circ}$, ΔH_r° is negative, so the equation can be written as follows:

$H_{2(g)} + O_{2(g)} \rightarrow H_2O_{(l)} \Delta H_r^{\circ} = -285.8\ \text{kJ}$

Similarly, the enthalpy of an endothermic reaction may be written as a reactant or as a separate positive term.

Practice

9. The reason dynamite releases a great amount of heat energy when it explodes is that the

A. products have more potential energy than the reactants in this endothermic reaction

B. reactants have more potential energy than the products in this endothermic reaction

C. products have more potential energy than the reactants in this exothermic reaction

D. reactants have more potential energy than the products in this exothermic reaction

12.2.2.2 *write thermochemical equations, expressing the energy change as a ΔH value or as a heat term in the equation*

THERMOCHEMICAL EQUATIONS

Chemical reaction equations can be written to include the enthalpy change. Endothermic reactions are written with the energy on the left side of the reaction equation. Exothermic reactions are written with the energy on the right side of the reaction equation. Consider the given examples:

Example

Consider the combustion of propane, exothermic, $\Delta H = -2\ 043.9$ kJ.

$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)} + 2\ 043.9$ kJ

Example

Write the equation that would describe the formation of dinitrogen tetroxide, endothermic, $\Delta H = +11.1$ kJ.

$N_{2(g)} + 2O_{2(g)} + 11.1\text{ kJ} \rightarrow N_2O_{4(g)}$

When energy is shown as part of the chemical reaction, the equation is called a thermochemical equation.

Practice

10. The thermochemical equation that represents the molar enthalpy of formation for ethanol is

A. $C_2H_5OH_{(l)} + 277.1\text{ kJ} \rightarrow 2C_{(s)} + 3H_{2(g)} + \frac{1}{2}O_{2(g)}$

B. $C_2H_5OH_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(g)} + 1\ 235.3\text{ kJ}$

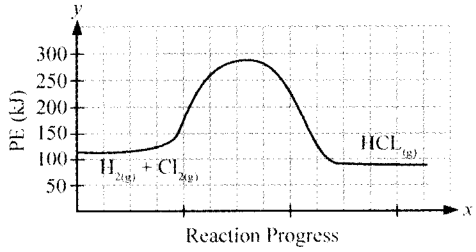
C. $2C_{(s)} + 3H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow C_2H_5OH_{(l)} + 277.1\text{ kJ}$

D. $C_2H_5OH_{(l)} + H_2O_{(l)} \rightleftharpoons C_2H_5O^-_{(aq)} + H_3O^+_{(aq)} K = 1.3 \times 10^{-16}$

Open Response

Use the following information to answer the next question.

The given graph shows the energy required to form $HCl_{(g)}$.



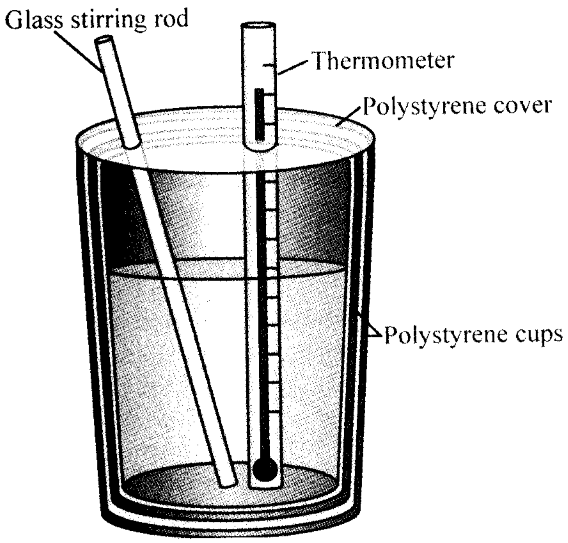
11. Use the given information to write the thermochemical equation for the decomposition of $HCl_{(g)}$.

12.2.2.3 *determine heat of reaction using a calorimeter, and use the data obtained to calculate the enthalpy change for a reaction*

CALORIMETRY

In a simple styrofoam cup calorimeter, energy changes in a system are measured by observing temperature changes in a measured volume of water in the cups. The energy lost by the process will be equal to the energy gained by the water, or the energy gained by the process will be equal to the energy lost by the water. One has to assume that solutions produced in a styrofoam cup calorimeter have the same density and specific heat capacity as water. One also has to assume that energy is not absorbed or released from the surroundings, including the calorimeter materials unless stated otherwise in the question.

Polystyrene calorimeter



To calculate the amount of energy absorbed or released, use $\Delta H = Q = mc\Delta t$, where m, c, and Δt all refer to water. To find the molar enthalpy for a particular substance in the process, use $\Delta H = n\Delta_r H$, where $n =$ moles of substance undergoing the change, and $\Delta_r H$ is the molar enthalpy for the process. The symbol "r" stands for the process. Symbols like "c" for combustion, "soln" for solution, and others are used. Consider the given example:

Example

A 2.50 g sample of NaOH is dissolved in 100.0 mL of water in a styrofoam cup calorimeter. The temperature of the water rose from 21.2°C to 28.5°C. Calculate the molar enthalpy of solution for NaOH.

$$n\Delta_{\text{soln}}H = mc\Delta t$$

$$n = \frac{2.50\ \text{g}}{40.00\ \text{g/mol}} = 0.0625\ \text{mol}$$

$$\Delta_{\text{soln}}H = \frac{mc\Delta t}{n}$$

$$= \frac{0.1000\ \text{kg} \times 4.19\ \text{kJ/kg}\cdot{}^\circ\text{C} \times (28.5 - 21.2^\circ\text{C})}{0.0625\ \text{mol}}$$

$$= -48.9\ \text{kJ/mol}$$

The answer is negative, since the temperature of the water increases and the dissolving process is exothermic.

In a bomb calorimeter, the reaction takes place in a container called a bomb inside a highly insulated calorimeter. All of the components—the thermometer, the stirrer, the bomb, and the container —are included in a constant called the heat capacity of the calorimeter (C). The heat capacity C has units kJ / °C. Calculations are similar to the given calculations. The new formula is $n\Delta_r H = C\Delta t$. The calorimeter's heat capacity includes the water and all the components, thus there is no need for a term describing mass.

Practice

Use the following information to answer the next question.

In a calorimeter containing 250 mL of water, 15 g of a substance was burned. After combustion, the temperature of the water increased by 33 °C. However, it was found later that the substance burned inside the calorimeter did not burn completely and 2 g of it remained. The density of water at the given temperature is 0.9973.

12. What is the enthalpy of the reaction?

A. −1 797 J

B. −2 074 J

C. −34 474 J

D. −34 567 J

Numerical Response

Use the following information to answer the next question.

A chemist heated 500 g of water in a calorimeter from 25.0 °C to 91.0 °C using 0.133 mol of ethanol.

13. If it is assumed that all the heat energy was absorbed by the water, the experimental molar enthalpy of combustion for ethanol was ___ MJ / mol.
(Record your **three-digit** answer.)

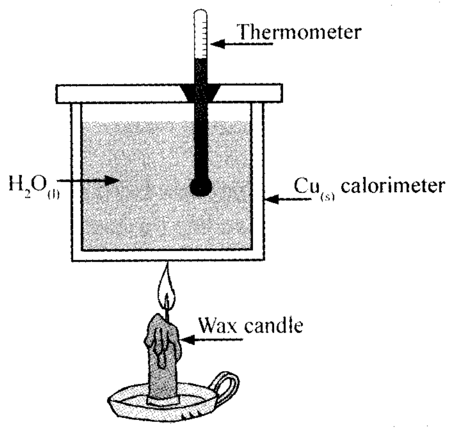
Numerical Response

14. A 24.6 g sample of molten copper at its melting point is lowered into a calorimeter containing 200 g of water. As soon as solidification is complete, the sample is quickly removed. The temperature of the water in the calorimeter rises from 11.23°C to 17.34°C.

In this experiment, the molar heat of solidification for copper is ____ kJ/mol.
(Record your **three-digit** answer.)

Open Response

Use the following information to answer the next question.



The given apparatus was used in an experiment to determine the molar enthalpy of the combustion of candle wax, $C_{25}H_{52(s)}$.

15. List all the measurements that must be taken in order to determine the molar enthalpy of combustion.

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12.2.2.4 *apply Hess's law to solve problems, including problems that involve data obtained through experimentation*

HESS'S LAW

The enthalpy change of a system depends only on the initial and final states of the system, and not on how the change took place. In other words, if a change takes place in one step and the same change takes place in two or more steps, the value of ΔH_r° is the same for both the single-step process and the multistep process. Such a quantity is called a **state function**.

Hess's law makes use of this property of ΔH. It states that for any multistep reaction, the enthalpy change for the overall reaction is equal to the sum of the enthalpy changes for each individual step. This can be used to determine the enthalpy change for a one-step reaction that is difficult to carry out, but the reaction can be easily done in a series of steps.

If the value of ΔH for a reaction is known, the ΔH for the reverse reaction has the same value, but with the opposite sign.

Example

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}\Delta H_r^{\circ} = -91.8 \text{ kJ}$$

$$2NH_{3(g)} \rightarrow N_{2(g)} + 3H_{2(g)}\Delta H_r^{\circ} = +91.8 \text{ kJ}$$

The equation, $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$ is the equation for the production of two moles of ammonia. The equation for the production of one mole of ammonia can be written as follows:

$$\frac{1}{2}N_{2(g)} + \frac{3}{2}H_{2(g)} \rightarrow NH_{3(g)}\Delta H_r^{\circ} = -45.9 \text{ kJ.}$$

When an equation is divided by 2, the value of the enthalpy is also divided by 2. Similarly, a thermochemical equation can be multiplied by any number, as long as the enthalpy change is multiplied by the same number.

Example

Use Hess's law to calculate the value of ΔH° for the reaction $2Fe_{(s)} + \frac{3}{2}O_{2(g)} \rightarrow Fe_2O_{3(s)}$, given the following thermochemical equations:

$$Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)} \quad \Delta H^{\circ} = -26.7 \text{ kJ}$$

$$CO_{(g)} + \frac{1}{2}O_{2(g)} \rightarrow CO_{2(g)}\Delta H^{\circ} = -383.0 \text{ kJ}$$

Solution

The equation for which the value of ΔH° is to be determined is called the "target equation".

The other equations can be numbered as follows:

1. $Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)}$ $\Delta H^{\circ} = -26.7$ kJ
2. $CO_{(g)} + \frac{1}{2}O_{2(g)} \rightarrow CO_{2(g)}\Delta H^{\circ} = -383.0$ kJ

The target equation shows the iron atoms on the left. This means that equation (1) needs to be reversed so that the iron atoms are on the left. The number of iron atoms in equation (1) is the same as the number of iron atoms in the target equation. This means that no further manipulation is required.

$$-1(1)2Fe_{(s)} + 3CO_{2(g)} \rightarrow Fe_2O_{3(s)} + 3CO_{(g)} \quad \Delta H^{\circ} = +26.7 \text{ kJ}$$

The -1 indicates that the equation has been reversed. The enthalpy term is also multiplied by -1.

The target equation shows oxygen molecules on the left and equation (2) also has the oxygen on the left. However, equation (2) has only $\frac{1}{2}$ mole of O_2 and $\frac{3}{2}$ mole of O_2 is needed for the target equation. Therefore, multiply equation (2) by 3.

$$3(2)3CO_{(g)} + \frac{3}{2}O_{2(g)} \rightarrow 3CO_{2(g)}\Delta H^{\circ} = -849.0 \text{ kJ}$$

Add the two equations together to determine the target equation.

$$-1(1)2Fe_{(s)} + 3CO_{2(g)} \rightarrow Fe_2O_{3(s)} + 3CO_{(g)} \quad \Delta H^{\circ} = +26.7 \text{ kJ}$$

$$3(2)3CO_{(g)} + \frac{3}{2}O_{2(g)} \rightarrow 3CO_{2(g)}\Delta H^{\circ} = -849.0 \text{ kJ}$$

$$2Fe_{(s)} + \frac{3}{2}O_{(g)} \rightarrow Fe_2O_{3(s)}\Delta H^{\circ} = -822.3 \text{ kJ}$$

The reaction represented by the target equation does not need to be carried out in order to determine the value of ΔH for the reaction as long as the reactions represented by equations (1) and (2) can be carried out.

Practice

Use the following information to answer the next question.

I. $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

$\Delta H_1 = -393.5$ kJ

II. $C_{(s)} + \frac{1}{2}O_{2(g)} \rightarrow CO_{(g)}$

$\Delta H_2 =$ unknown

$CO_{(g)} + \frac{1}{2}O_{2(g)} \rightarrow CO_{2(g)}$

$\Delta H_3 = -283.9$ kJ

16. Carbon dioxide can be formed by either the single-step process (**I**) or the two-step process (**II**) shown. What is the value of ΔH_2?

A. −677.4 kJ

B. −109.6 kJ

C. 109.6 kJ

D. 677.4 kJ

Open Response

17. Which of the following step reactions are needed in the calculation of the enthalpy of formation for aluminium chloride? Prove your answer.

TARGET: $2Al_{(s)} + 3Cl_{2(g)} \rightarrow 2AlCl_{3(s)}$

1. $2Al_{(s)} + 6HCl_{(aq)} \rightarrow 2AlCl_{3(aq)} + 3H_{2(g)}$
2. $AlCl_{3(s)} + H_2O_{(l)} \rightarrow AlCl_2OH_{(s)} + HCl_{(g)}$
3. $HCl_{(g)} \rightarrow HCl_{(aq)}$
4. $3H_{2(g)} + 3Cl_{2(g)} \rightarrow 6HCl_{(g)}$
5. $Al(OH)_{3(s)} + HCl_{(aq)} \rightarrow H_2O_{(l)} + AlCl_{3(aq)}$
6. $AlCl_{3(s)} \rightarrow AlCl_{3(aq)}$

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Numerical Response

Use the following information to answer the next question.

1. $N_{2(g)} + 2O_{2(g)} \rightarrow 2NO_{2(g)}$

$\Delta H = +67.7$ kJ

2. $N_{2(g)} + 2O_{2(g)} \rightarrow 2N_2O_{4(g)}$

$\Delta H = +9.7$ kJ

18. The absolute value of ΔH for $2NO_{2(g)} \rightarrow N_2O_{4(g)}$ is ____ kJ.
(Record your **three-digit** answer.)

Use the following information to answer the next question.

Carbon monoxide is formed by the incomplete oxidation of carbon according to the given reaction.
$2C + O_2 \rightarrow 2CO$. The enthalpy change for this reaction is
given as $\Delta H_1 = -110.5$ kJ / mol

When carbon monoxide is burned in a sufficient supply of oxygen, carbon dioxide is formed according to the given reaction.
$2CO + O_2 \rightarrow 2CO_2$ The enthalpy change for this reaction is
given as $\Delta H_2 = -283.9$ kJ / mol

19. What is the enthalpy of formation of two moles of carbon dioxide?

A. −394.4 kJ

B. −283.9 kJ

C. −173.4 kJ

D. −110.5 kJ

12.2.2.5 *calculate heat of reaction using tabulated enthalpies of formation*

STANDARD ENTHALPIES OF FORMATION

If the tabulated information (the heat energies, or the enthalpies of formation of all the species in the reaction) is available, the enthalpy change of a net reaction is the difference between the total enthalpies of formation of the products and reactants. In equation form, write

$$\Delta H_r^\circ = \sum\left(n\Delta H^\circ_{\text{f products}}\right) - \sum\left(n\Delta H^\circ_{\text{f reactants}}\right)$$

($\sum$ means "the sum of")

For example, the enthalpy of combustion of butane is calculated as follows:

$$C_4H_{10(g)} + \frac{13}{2}O_{2(g)} \rightarrow 4CO_{2(g)} + 5H_2O_{(g)}$$

$$\Delta H^\circ = \left[4 \text{ mol} \times H_f^\circ(CO_{2(g)})\right]$$
$$+\left[5 \text{ mol} \times H_f^\circ(H_2O_{(g)})\right] - \left[1 \text{ mol} \times H_f^\circ(C_4H_{10(g)})\right]$$
$$\Delta H^\circ = [4 \text{ mol} \times (-393.5 \text{ kJ/mol})]$$
$$+[5 \text{ mol} \times (-241.8 \text{ kJ/mol})]$$
$$-[1 \text{ mol} \times (-126.5 \text{ kJ/mol})]$$
$$\Delta H^\circ = -2\,656.5 \text{ kJ}$$

The enthalpy of combustion of butane is −2 656.5 kJ. (You could just as easily say the molar enthalpy of combustion of butane to give gaseous products is −2 656.5 kJ/mol.)

Practice

Use the following information to answer the next question.

Standard Heats of Formation

Substance	**ΔH_f°(kJ/mol)**
X	−22.5
Y	+78.3
Z	−54.8

Given $X + 3Y \rightleftharpoons 2Z + 2W$
$\Delta H = -562.0$ kJ

20. The standard molar heat of formation of substance W is

A. +442.0 kJ/mol

B. −120.0 kJ/mol

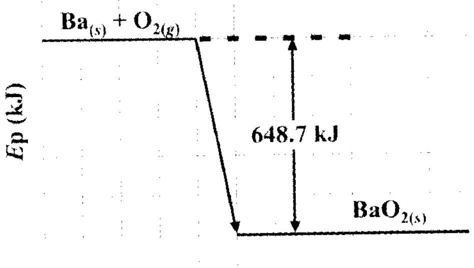
C. −240.0 kJ/mol

D. −451.4 kJ/mol

Numerical Response

Use the following information to answer the next question.

Heat of Formation for $BaO_{2(s)}$



Reaction coordinate

21. According to the information given in the data booklet, ΔH of reaction for $BaO_{(s)} + \frac{1}{2}O_{2(g)} \rightarrow BaO_{2(s)}$ is +/−____ kJ. (Record your **three-digit** answer.)

Open Response

Use the following information to answer the next question.

The combustion of sugar in a bomb calorimeter is similar to the oxidation of sugar in the body. A student ate three sugar cubes with masses of 6.84 g, 6.75 g, and 6.79 g, respectively.

22. Calculate the overall molar enthalpy of oxidation of sugar, $C_{12}H_{22}O_{11(s)}$, in the body.

12.2.2.6 *determine through experimentation a rate of reaction, and measure the effect on it of temperature, concentration, and catalysis*

REACTION RATE FACTORS

There are four main factors or conditions that can alter the chemical rates of different reactions. The amount of influence these factors have on reactions can be observed by manipulating that condition while observing a reaction. Observing the given properties of products and reactants allows conclusions to be drawn about the factors that determine the rate of each reaction.

CHANGE IN PRESSURE (FOR GASEOUS SYSTEMS ONLY)

Example

$PCl_{5(g)} \rightarrow PCl_{3(g)} + Cl_{2(g)}$

CHANGE IN MASS

Example

$2Li_{(s)} + 2H_2O_{(l)} \rightarrow H_{2(g)} + 2LiOH_{(aq)}$

CHANGE IN pH (FOR SYSTEMS INVOLVING AN ACID OR BASE)

Example

$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

CHANGE IN ELECTRICAL CONDUCTIVITY (FOR SYSTEMS PRODUCING IONS)

Example

$(CH_3)_3Br_{(aq)} + H_2O_{(l)} \rightarrow (CH_3)_3OH_{(aq)} + H^+_{(aq)} + Br^-_{(aq)}$

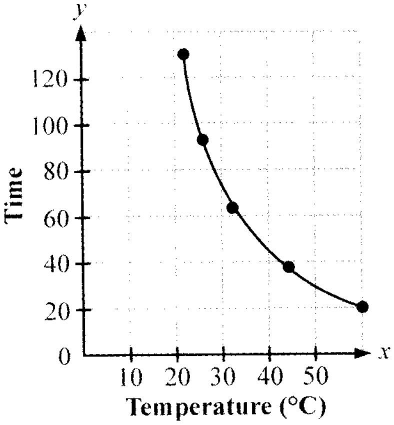
The observation of the previously listed four types of change in the reactions mentioned leads to the conclusion that there are four variables that will speed up or slow down a reaction:

- temperature
- surface area of reactants
- concentration of reactants
- presence of a catalyst

According to the collision model of atoms, increasing the surface area or concentration increases the number of molecular collisions as more reactant molecules are available to react. The addition of a catalyst decreases the activation energy required thereby, increasing the number of effective collisions. Finally, increasing the temperature gives more kinetic energy to reactant molecules, thus increasing both the number of collisions and the number of effective collisions. These changes are observable experimentally when one of the four factors that affect reaction rate are altered.

Practice

Use the following information to answer the next question.



23. According to the given graph, the rate of the reaction increases with a rise in temperature, because an increase in temperature

A. increases the kinetic energy of the reactants

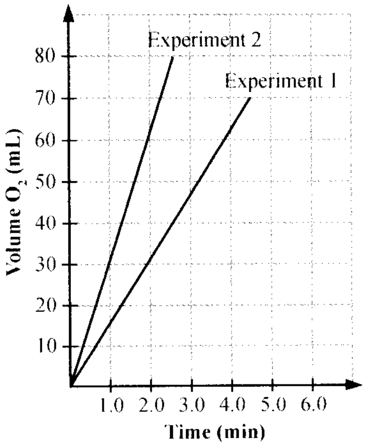
B. decreases the kinetic energy of the reactants

C. increases the potential energy of the reactants

D. decreases the potential energy of the reactants

Use the following information to answer the next multipart question.

24. The release of $O_{2(g)}$ resulting from the decomposition of bleach was measured in two different experiments. Data was collected and the following graph was drawn.



Part A

Open Response

Calculate the average rate of reaction for each experiment.

Part B

Open Response

Identify a variable from Experiment 1 and how it was changed to produce the different reaction rate for Experiment 2. Explain using collision theory.

12.2.3.1 *compare conventional and alternative sources of energy with respect to efficiency and environmental impact*

APPLICATIONS OF ENTHALPY AND RATES

Industrial uses of reaction rate and enthalpy change are widespread. Using enthalpy of reaction, alternate fuel sources can be compared in terms of their relative energy production to determine which is the most efficient with the least environmental impact. There are also a large variety of endothermic or exothermic reaction-related technologies, such as hot and cold packs and hydrogen rocket fuel.

Reaction rates are often controlled to optimize a desired effect. For example, food is cooled and metal products are galvanized to prevent decay and rusting, respectively. Other reactions, like the combustion of various fuels or the production of an industrial chemical are sped up to meet demand and increase product output. In this section, you will use principles learned in this unit as well as print and electronic sources to investigate the relationship between rate and enthalpy and their various industrial applications.

Practice

Use the following information to answer the next question.

The sun is the greatest reserve of energy relative to Earth. Solar energy can be used in a variety of activities like cooking food or running a solar-powered car.

25. In order to preserve the steadily depleting exhaustible sources of conventional energy, various non-conventional sources are being tapped.

The **largest** potential is possessed by

A. water power **B.** wind power

C. solar power **D.** bio mass

Open Response

26. In an effort to reduce the amount of fossil fuels burned to produce electricity, power companies are researching alternative methods of producing electricity.

Describe two alternative methods of producing electricity that use renewable resources. List two advantages and two disadvantages of each.

12.2.3.2 *describe examples of technologies that depend on exothermic or endothermic changes*

ENERGY CHANGE AND TECHNOLOGY

Hot packs and cold packs are applications of exothermic and endothermic changes, respectively. A hot pack contains the reactants of an exothermic reaction. The reactants are separated until the hot pack is needed, at which point in time the membrane separating the reactants is broken, and the reactants are allowed to mix. The resulting exothermic reaction loses heat to the surroundings. The reactants may be as simple as a salt and water.

A cold pack works by a similar process except that when the two reactants are mixed, the resulting reaction is an endothermic change.

Combustion reactions are exothermic. Most homes are heated by combustion reactions.

Refrigerators and air conditioners rely on an endothermic change to absorb heat from the surroundings, although the change is not a chemical change. A gas is compressed mechanically and then allowed to expand. The expansion is an endothermic process.

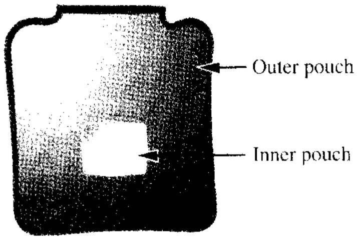
When your body becomes infected by a virus or bacterium, your temperature often increases due to exothermic biochemical changes.

There are numerous other examples of exothermic and endothermic processes that you can research.

Practice

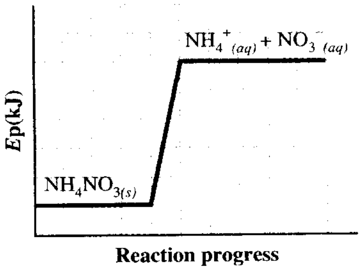
Use the following information to answer the next question.

Commercially available "cold packs" and "hot packs" contain an inner pouch of a solid ionic compound within an outer pouch containing water. When the inner pouch is broken, the solid dissolves in the water of the outer pouch. When ammonium nitrate dissolves, the water temperature decreases; whereas, when calcium chloride dissolves, the water temperature increases.

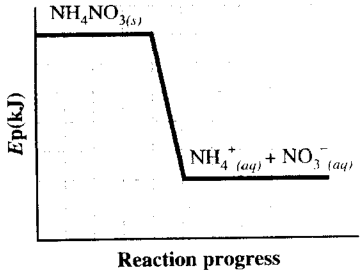


27. Which of the following diagrams represents the heat of solution for either a cold pack or a hot pack?

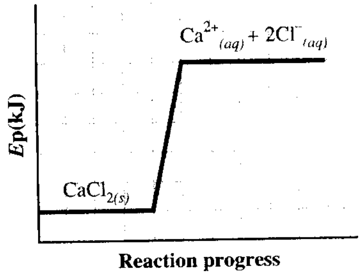
A.



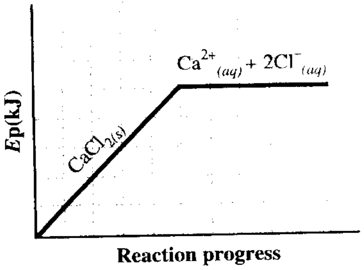
B.



C.



D.



12.2.3.3 *describe the use of catalysts in industry and in biochemical systems on the basis of information gathered from print and electronic sources*

USES OF CATALYSTS

Chemical processes used in industry are often too slow to be economical. Catalysts are used to increase the rate of production. These are referred to as **industrial catalysts**.

Biological systems, including your body, depend on catalysts to regulate the speed of the biochemical reactions in the system. These catalysts are called **enzymes** and are specific to one particular biochemical change. Enzymes are very large protein molecules.

RESEARCH IN CHEMISTRY

To learn more about industrial catalysts and enzymes, you may go use your library and search the internet. Whenever you research a topic on the internet, you should be aware of the pitfalls. Educational sites such as university sites and government agency sites are usually reliable but commercial sites such as those with the extension ".com" may not be as reliable. The primary goal of a commercial site is not education.

You should check the credentials of the author. Is this person qualified to write about the topic that you are researching? Is the publisher identified? Has the site been rated by a trustworthy organization? What is the purpose of the site? Is the site biased? Does it stereotype? Are the site's links working and are they trustworthy? Is the site documented? Is the site easy to use and to navigate? Is the information current? Check the date!

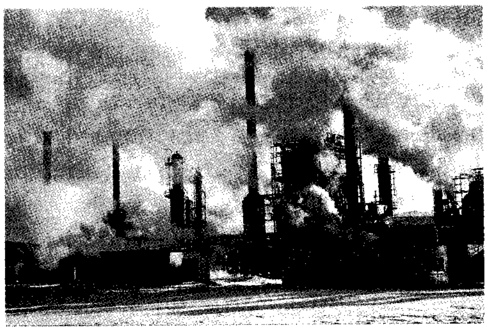
When you write a research paper, remember to give credit to any person or organization whose ideas you have used or to whose works you have referred.

Practice

Open Response

Use the following information to answer the next question.

Crude oil is a mixture of hydrocarbons, some of which are used in the production of gasoline. During the refining process the desired molecules are extracted through a type of distillation. However, the demand for gasoline is greater than what is naturally available. A process called catalytic cracking is often used to alter hydrocarbons in crude oil to increase fuel production. Below is a picture of a petrochemical refinery.



28. In a concise paragraph, outline the following:

- What is catalytic cracking and how does it work?
- What molecules are used as catalysts?
- Why is catalytic cracking important for engine performance?

12.2.3.4 *describe examples of slow chemical reactions, rapid reactions, and reactions whose rates can be controlled*

DIFFERENT RATES OF REACTIONS

Some chemical processes need to be fast to be effective, like the burning of rocket fuel or the detonation of explosives or fireworks. Other reactions need to be slow to be of use. For example, some antibiotics need to be released slowly into the bloodstream so that they are in contact with the invading bacteria for the maximum length of time or air fresheners need to release their scent slowly over a period of time. However, if an industrial process is too fast or too slow, then the speed of the reaction needs to be modified. If it is too slow, then it might not be economical. If it is too fast, then it could get out of control or generate too much heat.

Catalysts are used to speed up a reaction. Substances that slow down a reaction are called **inhibitors**. There are other means of controlling the rates of reactions. You could change the temperature, change the concentration, or remove one product as it is formed. In the kitchen, such methods are used all the time. When food is placed into the refrigerator, its temperature is lowered to slow down the reactions that cause food to spoil. When food is cooked, the high temperature speeds up the processes that make the food tender or produce flavour-enhancing chemicals.

Through research, you can find many examples of reactions that should proceed rapidly and those that should proceed slowly.

Practice

29. The formation of iron rust is accelerated by the presence of

A. N_2 in the air

B. acid in water

C. galvanizing

D. alkali in water

SOLUTIONS–ENERGY CHANGES AND RATES OF REACTION

1. D	**7. OR**	**14. −13.2**	**21. 100.7**	**27. A**
2. A	**8. C**	**15. OR**	**22. OR**	**28. OR**
3. OR	**9. D**	**16. B**	**23. A**	**29. B**
4. D	**10. C**	**17. OR**	**24.** Part A- **OR**	
5. Part A- **OR**	**11. OR**	**18. 58.0**	Part B- **OR**	
Part B- **OR**	**12. C**	**19. A**	**25. C**	
6. A	**13. 1.04**	**20. B**	**26. OR**	

1. D

All the processes listed involve making bonds or breaking bonds.

Phase changes involve making or breaking intermolecular bonds.

Chemical changes involve making and breaking intramolecular bonds.

Nuclear changes involve making and breaking intranuclear bonds.

Since the magnitude of the bond energies involved decreases in the order:

intranuclear > intramolecular > intermolecular, the energy changes arranged in order of decreasing magnitude are nuclear fusion, chemical, and phase.

2. A

Hess's law implies that as long as the initial reactants and final products of a chemical or physical change are the same, the number of steps used does not change the net enthalpy change of that process. Hence, the enthalpy changes for both processes are the same.

3. Open Response

No, the correct ΔH value for this reaction cannot be found since only the enthalpy of formation of water vapour is given. The enthalpy of formation of water liquid is needed to fulfill Hess's law.

Example: $C_2H_{4(g)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + \mathbf{2H_2O_{(l)}}$

$\Delta H = -1\ 411$ kJ

$2CO_{2(g)} + 3H_2O_{(l)} \rightarrow C_2H_{6(g)} + \frac{7}{2}O_{2(g)}$

$\Delta H = +1\ 560$ kJ

$H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow \mathbf{H_2O_{(g)}} \Delta H = -241.8$ kJ

Water vapour and liquid water do not cancel out.

4. D

Except for zero order rates, all reactions start out fast and end slow as the reactant concentration(s) decreases. This is precisely what is illustrated in graph D. **Note:** graph A resembles a distribution of collision/particle energies; graph C resembles the energy profile of an endothermic reaction; graph B is quite unlikely.

5. Part A – Open Response

The average rate of hydrolysis of butyl chloride,

$$r = \left\{\frac{[C_4H_9Cl]_{400} - [C_4H_9Cl]_{300}}{t_2 - t_1}\right\}$$

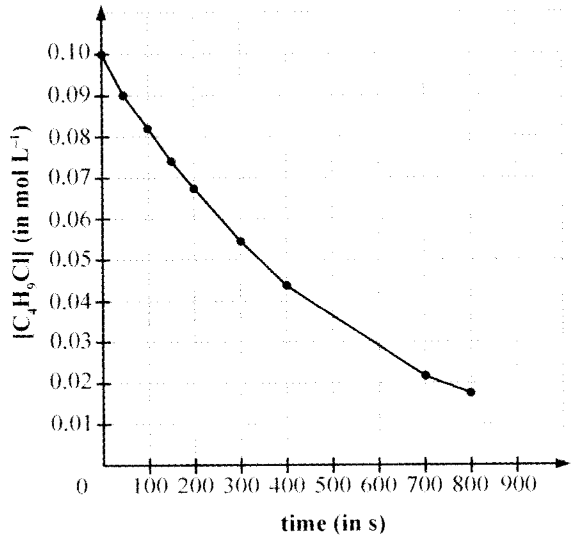
$$= -\left\{\frac{0.0439 - 0.0549}{100}\right\}$$

$$= \frac{0.0110}{100}$$

$$r = 1.1 \times 10^{-4}\text{mol/L}\cdot\text{s}$$

Part B – Open Response

The instantaneous rate of a reaction can be found by plotting a graph of the reaction as



$$r_{\text{in s}} \text{ at } 600 \text{ s} = \frac{0.0124 - 0.0400}{800 - 400}$$

$$= \frac{0.0276}{400}$$

$$= 6.9 \times 10^{-5}\text{mol/L}\cdot\text{s}$$

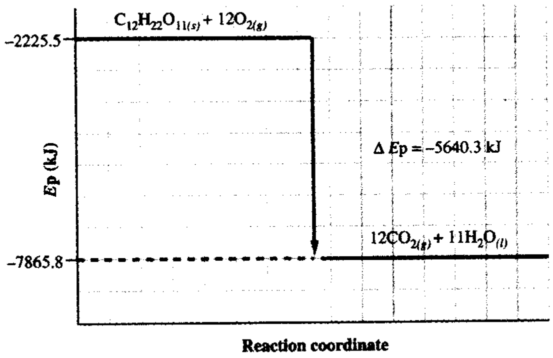
6. A

The E variable represents activation energy, which is the energy required for the reactant to reach its activated state. Activation energy behaves as an energy barrier between the reactants and the activated complex.

7. Open Response

E_p diagram for $C_{12}H_{22}O_{11(s)}$

Ep diagram for $C_{12}H_{22}O_{11(s)}$



8. C

The rate law of the reaction is not deduced directly from the stoichiometry of the balanced chemical equation. If it had been, then the expression should be

rate = $k[H_2][Br_2]$

This suggests that the chemical reaction has occurred by a mechanism that involves more than one step. The reaction occurs by the given mechanism.

i. $Br_2 \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} 2Br$ (fast)

ii. $Br + H_2 \xrightarrow{k_2} HBr + H$ (slow)

iii. $H + Br_2 \xrightarrow{k_3} HBr + Br$

And the rate law has been deduced by placing the values of $[H_2]$ and [Br] in the rate expression given by the slowest step, that is

$$\frac{1}{2}\frac{d[HBr]}{dt} = k_2[H_2][Br]$$

$$[Br] = \left(\frac{k_1}{k_{-1}}\right)^{\frac{1}{2}}[Br_2]^{\frac{1}{2}} \text{ (from step } i\text{)}$$

$$\Rightarrow \text{rate} = k_2\left(\frac{k_1}{k_{-1}}\right)^{\frac{1}{2}}[H_2][Br_2]^{\frac{1}{2}}$$

Distractor Rationale

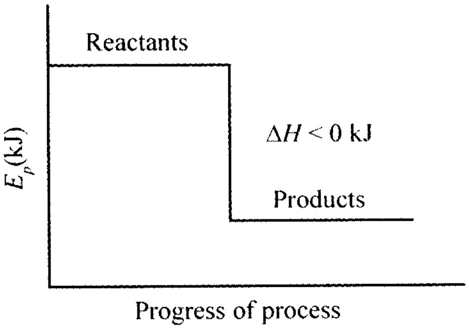
A,B. Only the rate law of elementary reactions can be deduced from the balanced chemical equations. Elementary reactions are those reactions that occur in one-step.

D. The rate expression does not contain [HBr] term. Hence, it is independent of concentration of HBr.

The correct answer is C.

9. D

A reaction that loses or releases energy is said to be exothermic. The energy profile of such a reaction is as follows:



The reactants have more E_p than do the products.

10. C

The $H^\circ_{f(C_2H_5OH_{(l)})} = -277.1$ / kJ.

The formation reaction of ethanol is

$$2C_{(s)} + 3H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow C_2H_5OH_{(l)}$$

The negative sign in the formation enthalpy signifies that energy is lost (or *produced*) when ethanol forms from its elements.

$$2C_{(s)} + 3H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow C_2H_5OH_{(l)} + 277.1 \text{ kJ}$$

Note: Alternative A represents an endothermic reaction, alternative B is the combustion of ethanol, and alternative D represents the "acidity" of ethanol in water, which is negligible given the extremely small K value.

11. Open Response

$$2HCl_{(g)} + 50 \text{ kJ} \rightarrow H_{2(g)} + Cl_{2(g)}$$

12. C

The heat released during a process can be calculated by using the given formula.

$Q = mc\Delta T$
where Q = heat energy
m = mass
c = specific heat capacity
ΔT = temperature

According to the information given in the question,

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

$m = (0.9973 \times 250)$ g = 249.325 g
$\Delta T = 33$°C
The specific heat capacity of water $c = 4.19$ J / g / °C
$\therefore Q = (249.325 \times 4.19 \times 33)$ J
$= 34\,474.17$ J

Since $H = -Q$, the enthalpy of the reaction is $-34\,474$ J.

13. 1.04

Rearrange the formula $nH = mc\Delta t$ to give H:

$$H = \frac{mc\Delta t}{n}$$

where H = molar enthalpy (units kJ / mol), m = mass (unit kg), c = specific heat capacity (unit J / g°C), Δt = temperature (unit°C) and n = number of moles (unit mol)

Replace the known variables with the given values:

$$H = \frac{0.500\ \text{kg} \times \frac{4.19\ \text{kJ}}{\text{kg} \cdot {}^\circ\text{C}} \times (91.0^\circ\text{C} - 25.0^\circ\text{C})}{0.133\ \text{mol}}$$

$$= 1.04 \times 10^3\ \text{kJ} / \text{mol}$$

10^3 kJ = 1 MJ

$\therefore 1.04 \times 10^3\ \text{kJ} / \text{mol} = 1.04\ \text{MJ} / \text{mol}$

The experimental molar enthalpy of combustion for ethanol was 1.04 MJ / mol. The value is negative because the reaction was exothermic. The heat lost by the reaction warmed the water.

14. –13.2

The enthalpy of solidification of the copper is equivalent to the kinetic energy change of the calorimeter.

$$n_{\text{Cu}} H_{\text{solidification}} = (mc\Delta t)_{\text{calorimeter}}$$

$$H_{\text{solidification}} = \frac{(mc\Delta t)_{\text{calorimeter}}}{n_{\text{Cu}}}$$

$$= \frac{(0.200\ \text{kg})\left(4.19 \frac{\text{kJ}}{\text{kg}^\circ\text{C}}\right)(17.34 - 11.23)^\circ\text{C}}{24.6\ \text{g} \Big/ 63.55 \frac{\text{g}}{\text{mol}}}$$

$$= 13.2 \frac{\text{kJ}}{\text{mol}}$$

The molar enthalpy/heat of solidification of copper is 13.2 kJ/mol.

15. Open Response

Measurements

- mass or volume of water
- mass of $Cu_{(s)}$ calorimeter
- initial temperature of water
- highest temperature of water
- initial mass of candle
- final mass of candle

16. B

$$\Delta H_1 = \Delta H_2 + \Delta H_3$$

$$-393.5\ \text{kJ} = \Delta H_2 + (-283.9\ \text{kJ})$$

$$\Delta H_2 = -109.6\ \text{kJ}$$

The value of ΔH_2 is -109.6 kJ

17. Open Response

Equations 1, 3, 4, and 6 are needed.

Proof:

1. $2Al_{(s)} + 6HCl_{(aq)} \rightarrow 2AlCl_{3(aq)} + 3H_{2(g)}$
3. $6HCl_{(g)} \rightarrow 6HCl_{(aq)}$
4. $3H_{2(g)} + 3Cl_{2(g)} \rightarrow 6HCl_{(g)}$
6. $2AlCl_{3(aq)} \rightarrow 2AlCl_{3(s)}$

Yields: $2Al_{(s)} + 3Cl_{2(g)} \rightarrow 2AlCl_{3(s)}$

18. 58.0

Follow these steps to make the equations add to the desired equation:

Reverse equation (1).
Add equation (2) to the first equation.

$$\begin{array}{r} 2\,NO_{2(g)} \longrightarrow \cancel{N_{2(g)}} + \cancel{2O_{2(g)}} \\ \cancel{N_{2(g)}} + \cancel{2O_{2(g)}} \longrightarrow N_2O_{4(g)} \\ \hline 2NO_{2(g)} \longrightarrow N_2O_{4(g)} \end{array}$$

$$\begin{array}{l} \Delta H = -67.7\ \text{kJ} \\ \underline{\Delta H = +9.7\ \text{kJ}} \\ \Delta H = -58.0\ \text{kJ} \end{array}$$

The absolute net change in enthalpy is 58.0 kJ.

19. A

Hess' law states that the change in enthalpy during a chemical process is the same whether the process takes place in one or several steps.

The formation of carbon dioxide from carbon and oxygen can be assumed to occur in two steps.

$2C + O_2 \rightarrow 2CO$

$2CO + O_2 \rightarrow 2CO_2$

Adding the two reactions:

$2C + O_2 + 2CO + O_2 \rightarrow 2CO + 2CO_2$

$2C + 2O_2 \rightarrow 2CO_2$

Thus, the enthalpy change in this reaction is given as

$\Delta H = \Delta H_1 + \Delta H_2 = (-110.5) + (-283.9)$

$= -110.5 - 283.9 = -394.4$ kJ

20. B

Use the short form of Hess's law:

$$\Delta H^{\circ}_{\text{net}} = \sum nH^{\circ}_{\text{f (products)}} \quad \sum nH^{\circ}_{\text{f (reactants)}}$$

Substitute to get

$$-562.0\ \text{kJ} = \left[2\ \text{mol} \times \left(\frac{-54.8\ \text{kJ}}{\text{mol}}\right) + 2\ \text{mol} \times H^{\circ}_{\text{f (W)}}\right]$$

$$-\left[1\ \text{mol} \times \left(\frac{-22.5\ \text{kJ}}{\text{mol}}\right)\right] + 3\ \text{mol} \times \left(\frac{+78.3\ \text{kJ}}{\text{mol}}\right)$$

Rearranging this equation gives you

$$2\ \text{mol} \times H^{\circ}_{\text{f (W)}} = -562.0\ \text{kJ} + 212.4\ \text{kJ} + 109.6\ \text{kJ}$$

$$2\ \text{mol} \times H^{\circ}_{\text{f (W)}} = 240.0\ \text{kJ}$$

$$H^{\circ}_{\text{f (W)}} = -120.0\ \text{kJ/mol}$$

The standard molar heat/enthalpy of formation of substance W is 120.0 kJ/mol.

21. 100.7

From the graph, $H^{\circ}_{\text{f}(\text{BaO}_{2(\text{s})})} = -648.7$ kJ/mol.

From the Chemistry 12 Data Booklet, Standard Molar Enthalpies of Formation,

$H^{\circ}_{\text{f}(\text{BaO}_{(\text{s})})} = -548.0$ kJ/mol.

The enthalpy change for the reaction given

$\text{BaO}_{(\text{s})} + \frac{1}{2}\text{O}_{2(\text{g})} \rightarrow \text{BaO}_{2(\text{s})}$ is calculated using

Hess's law $\Delta H^{\circ}_{\text{net}} = \sum nH^{\circ}_{\text{f(products)}} - \sum nH^{\circ}_{\text{f(reactants)}}$

$$= \left(1\ \text{mol} \times \left\{-648.7\frac{\text{kJ}}{\text{mol}}\right\}\right) - \left(1\ \text{mol} \times \left\{-548.0\frac{\text{kJ}}{\text{mol}}\right\}\right)$$

$$= -100.7\ \text{kJ}$$

The enthalpy change of the reaction is −100.7 kJ.

22. Open Response

$$\text{C}_{12}\text{H}_{22}\text{O}_{11(\text{s})} + 12\text{O}_{2(\text{g})} \rightarrow 12\text{CO}_{2(\text{g})} + 11\text{H}_2\text{O}_{(\text{l})}$$

$$\Delta E_p = 12(-393.5\ \text{kJ}) + 11(-285.8\ \text{kJ}) - (-2\,226.1\ \text{kJ})$$

The molar enthalpy for the oxidation of sucrose is −5 639.7 kJ/mol.

23. A

In the given reaction, the increase in temperature accelerates the movement of the reactants. As a result, the collision rate of the reactant molecules increases. Therefore, the kinetic energy of the reacting particles increases, leading to an increase in the reaction rate.

24. Part A – Open Response

$$\text{Experiment 1: } \frac{60\ \text{mL}}{2.5\ \text{min}} = 24\ \text{mL/min}$$

$$\text{Experiment 2: } \frac{50\ \text{mL}}{4.5\ \text{min}} = 11\ \text{mL/min}$$

Part B – Open Response

Variable/Change	Explanation
Temperature is decreased.	Lower fraction of effective collisions.

OR

Concentration of reactants was decreased.	Fewer collisions.

25. C

The sun is the greatest reserve of energy relative to Earth. Solar energy can be used in a variety of activities, like cooking food or running a solar-powered car.

26. Open Response

Any two of the following answers:

1. Solar panels capture solar energy and convert it into electricity.

Advantages	Disadvantages
• minimal costs to operate (little maintenance) • uses a renewable resource • no pollution	• expensive • takes a lot of space • fragile • must have devices to store electrical energy • production can only happen during the daytime.

2. Wind-driven electricity production (windmills) capture the wind's energy to drive a turbine that will drive a generator to produce electricity.

Advantages	Disadvantages
• simple to operate • uses a renewable resource • no pollution	• takes a lot of space • production can only happen during windy periods

3. Using the movement of tides to turn turbine blades that will turn a generator to produce electricity.

Advantages	Disadvantages
• tides change water level twice a day • uses a renewable resource • no pollution	• can be difficult to find a good location to utilize this movement • not as efficient as constantly moving water • need to devise a system that can capture the energy as tides rise and fall

4. Geothermal energy can be captured to produce electricity. In places where there are cracks in Earth's crust, heat from the inner parts of Earth causes geysers and steam to rise.

Advantages	Disadvantages
• uses a renewable resource • produces a constant source of energy • no pollution	• can only be used in certain places on Earth • sometimes water is high in sulfur, which can cause health problems

5. Hydroelectric power can be captured as kinetic energy from moving water and be converted into energy.

Advantages	Disadvantages
• uses a renewable resource • no pollution	• causes a major change and possible destruction of ecosystems when dams are built • may cause loss of homes when dams are built

27. A

The products of an endothermic process have more potential energy than the reactants, while the products of an exothermic process have less potential energy than the reactants. Remember that the dissolving of $CaCl_{2(s)}$ is exothermic.

28. Open Response

Catalytic cracking involves the breaking of larger hydrocarbons into smaller ones by catalysts such as alumina (Al_2O_3) and silica (SiO_2). The hydrocarbon mixture is vapourized and passed over the heated catalysts. Carbon-carbon bonds are broken, leaving lighter alkane molecules.

For an engine to function optimally, the gasoline should only burn when ignited by the spark plug. Uncontrolled combustion can cause a shock wave through the engine which results in an engine "knock" or clanging sound. Engine "knock" can cause serious damage to the engine of a car. Engine "knock" is decreased using isomers of octane as fuel which can be created through catalytic cracking.

29. B

The formation of iron rust is accelerated by the presence of an oxidizing agent. An acid in aqueous solution would provide an excess of either positive hydronium $\left(H_3O^+\right)$ ions or hydrogen $\left(H^+\right)$ ions, leading to the reduction half-reaction provided in the table of standard electrode potentials in the appendices.

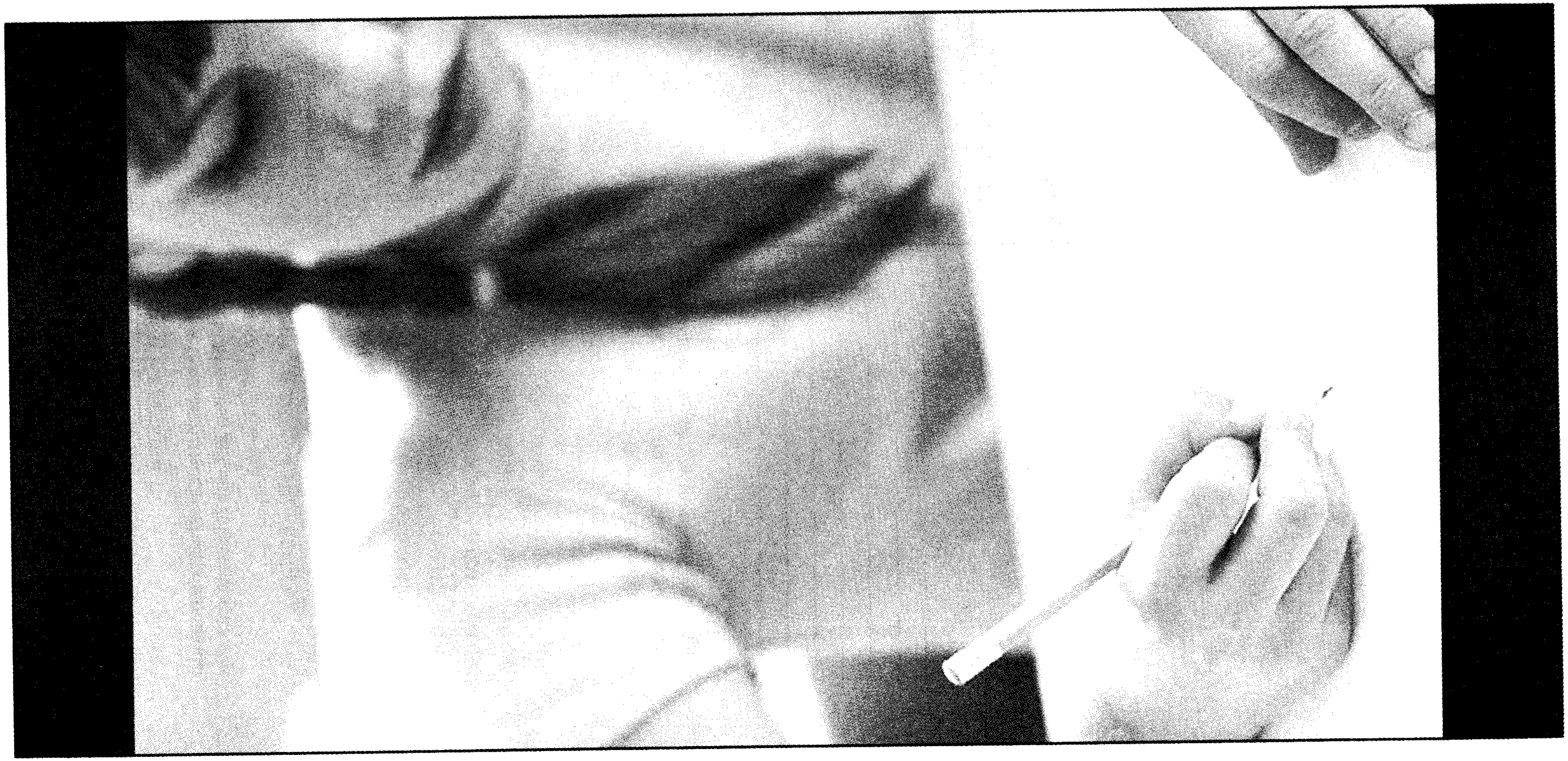
$O_{2(g)} + 4H^+ + 4e^- \rightleftharpoons 2H_2O_{(l)}$

The electrical potential of this half reaction provides more energy for oxidation than either nitrogen gas in the air (alternative A), or an alkaline solution (alternative D). Galvanizing actually prevents rusting, so alternative C is incorrect.

The correct response is B.

Carbon dioxide (CO_2) and water (H_2O) react to form carbonic acid in the air. Carbonic acid reacts with iron to form the corresponding carbonate that further decomposes to produce rust.

Self Test



1. The reason that nuclear reactions involve greater enthalpy changes than chemical reactions is that

A. nuclear forces are stronger than chemical bonds

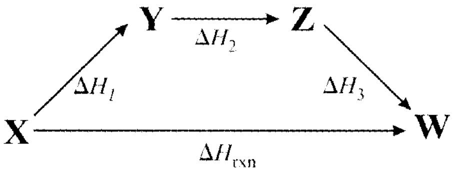
B. nuclear explosions are more energetic than chemical explosions

C. nuclear reactions involve different elements than chemical reactions

D. nuclear reactions involve elements with larger atomic numbers than chemical reactions

Use the following information to answer the next question.

The reaction X → W can proceed directly or through a series of steps, as shown in the given diagram.



2. This diagram illustrates

A. the law of conservation of mass

B. an endothermic reaction

C. an exothermic reaction

D. Hess's law

Open Response

3. In a reaction between nitrogen monoxide and hydrogen gas, the following data was recorded:

$$2NO_{(g)} + 2H_{2(g)} \rightarrow N_{2(g)} + 2H_2O_{(g)}$$

Trial	[NO] mol/L	[H$_2$] mol/L	Rate mol/L· s
1	0.001	0.004	0.002
2	0.002	0.004	0.008
3	0.003	0.004	0.018
4	0.004	0.001	0.008
5	0.004	0.002	0.016
6	0.004	0.003	0.024

Use this data to determine the rate order of the given reaction and give the rate law equation for the reaction between $NO_{(g)}$ and $H_{2(g)}$.

Open Response

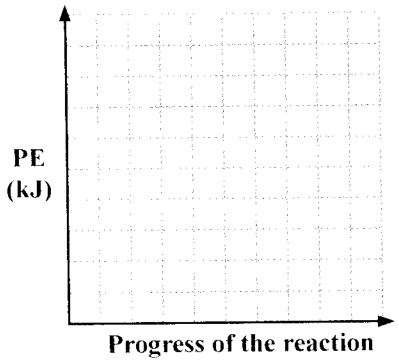
Use the following information to answer the next question.

Consider the following reaction mechanism:

Step 1 $NO_{(g)}+O_{2(g)} \rightarrow NO_{3(g)}$slow

Step 2 $NO_{3(g)}+NO_{(g)} \rightarrow 2NO_{2(g)}$

4. The overall reaction is exothermic. Sketch a PE diagram on the given axes to describe the energy changes that occur as the reaction takes place.



Use the following information to answer the next question.

The rate expression for the reaction $NO_{2(g)} + CO_{(g)} \rightarrow CO_{2(g)} + NO_{(g)}$ is $r = k[NO_2]^2$.

5. How many molecules of $CO_{(g)}$ are involved in the rate-determining step?

A. Zero **B.** One

C. Two **D.** Three

Use the following information to answer the next question.

$FeCl_{3(aq)} + 3NaOH_{(aq)} \rightarrow$
$Fe(OH)_{3(aq)} + 3NaCl_{(aq)}$

6. The only problem with the given equation is that

A. there is no mass balance in the equation

B. heat change during the reaction is not shown

C. one of the formulae of the compounds is wrong

D. physical states of the reactants and products are not shown

Numerical Response

Use the following information to answer the next question.

Cold packs are used to treat sprains and bruises.
A chemical commonly used in cold packs is ammonium nitrate, $NH_4NO_{3(s)}$, which can produce a cooling effect.
A student designed a calorimetry experiment to determine the energy change for the dissolving of ammonium nitrate and recorded the following results:

heat capacity of calorimeter and water = 228 J / °C

initial temperature of water = 21.6°C

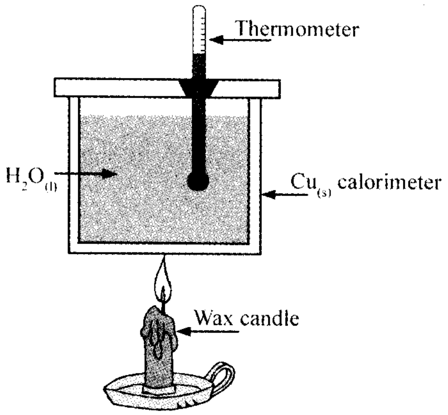
final temperature of water = 16.4°C

mass of ammonium nitrate = 0.250 g

7. The amount of energy involved in this change is ____ kJ.
(Record your **three-digit** answer.)

Open Response

Use the following information to answer the next question.



The above apparatus was used to determine experimentally the molar enthalpy of the combustion of candle wax, $C_{25}H_{52(s)}$.

8. Suggest two improvements to the experimental design.

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Numerical Response

Use the following information to answer the next question.

Ethanol is the alcohol found in beer, wine, and whiskey. In the production of ethanol, the starch in barley, grapes, or corn reacts to form glucose in the presence of enzymes. During the fermentation process, yeast is added to the glucose.

The yeast contains enzymes that act as biological catalysts in the reaction of glucose into ethanol and carbon dioxide. The reaction is represented by the equation:

$C_6H_{12}O_{6(s)} \rightarrow 2CO_{2(g)} + 2C_2H_5OH_{(l)}$

9. The molar enthalpy of reaction for the fermentation of glucose is –____ kJ/mol. (Record your **three-digit** answer.)

Open Response

Use the following information to answer the next question.

The uranium-235 isotope is used as a fuel in some nuclear power plants. This isotope is used to enrich natural uranium ore. Prior to the enrichment process, the uranium ore, $UO_{2(s)}$, is converted to $UF_{6(s)}$. This conversion is represented by the following sequential equations:

Equation I

$UO_{2(s)} + 4HF_{(g)} \rightarrow UF_{4(s)} + 2H_2O_{(g)}$

Equation II

$UF_{4(s)} + F_{2(g)} \rightarrow UF_{6(s)}$

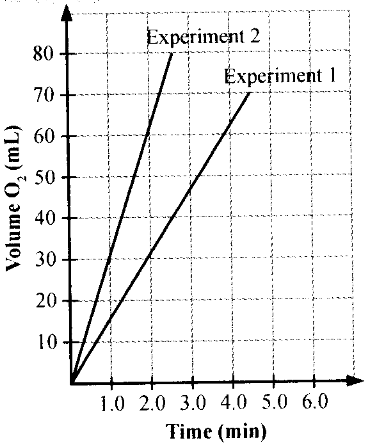
10. Use molar heats of formation to calculate the amount of heat energy involved in producing 2.00 Mg of $UF_{6(s)}$ from natural uranium ore, $UO_{2(s)}$.

Evaluate the use of nuclear energy for the generation of electricity.

Open Response

Use the following information to answer the next question.

The release of $O_{2(g)}$ resulting from the decomposition of bleach was measured in two different experiments. Data was collected and the following graph was drawn.



11. Identify a variable from Experiment 1 and how it was changed to produce the different reaction rate for Experiment 2. Explain using collision theory.

12. A difference between the generation of electricity in a coal-burning plant and that from a nuclear, wind, or solar plant is that the coal-burning plant

A. relies on a cheaper source of energy

B. makes use of renewable energy

C. produces much more electricity

D. generates greenhouse gases

Use the following information to answer the next question.

Cold packs are used to treat sprains and bruises. A chemical commonly used in cold packs is ammonium nitrate, $NH_4NO_{3(s)}$, which can produce a cooling effect.

13. The change that occurs in this cold pack is an

A. endothermic change, which results in an increase in temperature

B. exothermic change, which results in an increase in temperature

C. endothermic change, which results in a decrease in temperature

D. exothermic change, which results in a decrease in temperature

Use the following information to answer the next question.

Vanadium oxide (V_2O_5) is used to manufacture sulfuric acid (H_2SO_4) through a contact process. This is the industrial method of preparing H_2SO_4.

14. What role does vanadium oxide likely play in the production of sulfuric acid?

A. Spectator **B.** Catalyst

C. Reactant **D.** Solvent

15. Which of the following reactions has the **most** controlled rate?

A. Food decay

B. Rusting of metal

C. Explosion of TNT

D. Cellular respiration

Use the following information to answer the next question.

Many insects and small animals have unique defense systems. Bombardier beetles fight off predators with a hot chemical spray. This spray consists of solutions of hydroquinone ($C_6H_4(OH)_{2(aq)}$), hydrogen peroxide $H_2O_{2(aq)}$, and enzymes that are secreted by the beetles' glands.

Reaction Equations Related to Spray Formation

I. $2H_2O_{(l)} + O_{2(g)} \rightarrow 2H_2O_{2(aq)}$
$\Delta H = +189.2$ kJ

II. $H_2O_{(l)} \rightarrow H_{2(g)} + \frac{1}{2}O_{2(g)}$
$\Delta H = +285.8$ kJ

III. $C_6H_4(OH)_{2(aq)}$
$\rightarrow C_6H_4O_{2(aq)} + H_{2(g)}$
$\Delta H = +177.0$ kJ

A chemical reaction that occurs in order to produce the hot chemical spray can be represented by the equation

$$\underset{\text{hydroquinone}}{C_6H_4(OH)_{2(aq)}} + H_2O_{2(aq)} \rightarrow \underset{\text{quinone}}{C_6H_4O_{2(aq)}} + 2H_2O_{(l)}$$

16. The heat of reaction for the production of this hot chemical spray is

A. −489.2 kJ

B. −203.4 kJ

C. −82.4 kJ

D. +12.2 kJ

Use the following information to answer the next question.

The reactions involved in the formation of sulfur dioxide and sulfur trioxide are:

$S + O_2 \rightarrow SO_2 \quad (\Delta H_1 = -297\ \text{kJ})$

$S + \frac{3}{2}O_2 \rightarrow SO_3 (\Delta H_2 = -396\ \text{kJ})$

17. What is the enthalpy of the formation of sulfur trioxide from sulfur dioxide?

A. –49 kJ **B.** –99 kJ

C. –396 kJ **D.** –693 kJ

SOLUTIONS

1. A	**6. B**	**11. OR**	**16. B**
2. D	**7. 1.19**	**12. D**	**17. B**
3. OR	**8. OR**	**13. C**	
4. OR	**9. –69.1**	**14. B**	
5. A	**10. OR**	**15. D**	

1. A

All nuclear and chemical processes involve the making and breaking of bonds. The net enthalpy change of a process is the difference between the energy absorbed to break bonds in the reactants and the energy released to make bonds in the products. Nuclear reactions are accompanied by much greater enthalpy changes than chemical reactions simply because the bonds made and broken in nuclear processes are so much stronger.

2. D

Hess's law might otherwise be interpreted as "no matter what reaction pathway is broken between reactant X and product W, the enthalpy change for the reaction is the difference in formation enthalpies of X and W!" Alternative A has no bearing on thermodynamics; while alternatives B and C cannot be derived from the information given (it is not an enthalpy diagram). Put another way, if the route X → Y → Z → W was the catalysed alternative to direct reaction, X → W, the enthalpy change would be the same!

3. Open Response

When $[H_2]$ is constant:
When the trial [NO] rate is divided by two, the relationship between the concentration and the rate is shown in the table below.

[NO] mol/L	Rate
0.001	0.001
0.002	0.004
0.003	0.009

Looking at the relationship between the numbers, it can be seen that it is an exponential one, since $1^2 = 1$, $2^2 = 4$, and $3^2 = 9$. In other words, $r = [NO] \times [NO] \times 2$, or $r = 2[NO]^2$.
When [NO] is constant:
Dividing the trial $[H_2]$ rate by the concentration in each trial results in a constant, so the relationship is a direct one.

$$\frac{1}{8} = 8,\ \frac{2}{16} = 8,\ \frac{3}{24} = 8$$

In other words, $r = 8[H_2]$

Therefore the rate law is $r = k[NO]^2[H_2]$

To find the total rate order, add the exponents of the reactants: 2 + 1 = 3
The total rate order of the reaction is 3.

4. Open Response

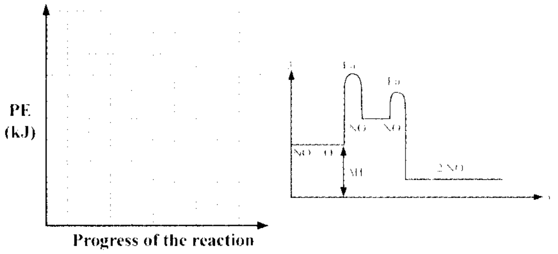


Diagram should show two peaks, one for each step. The first peak should be higher as it represents the first step which is slower and therefore has a higher activation energy to overcome. Reactants must have more energy than the products since it is an exothermic reaction.

5. A

CO is not present in the rate expression, so zero molecules of $CO_{(g)}$ are involved in the rate-determining step.

6. B

The given equation is not a thermochemical equation. In other words, the heat change is not shown.

7. 1.19

This question only asks for the energy involved, **not** the molar enthalpy of the reaction. Since the heat capacity of the calorimeter is quoted, the enthalpy change can be calculated:

$$\begin{aligned} \Delta H &= c\Delta t = 228\frac{\text{J}}{^\circ\text{C}} \times (21.6 - 16.4)^\circ\text{C} \\ &= 228\ \text{J}/^\circ\text{C} \times 5.2^\circ\text{C} \\ &= 1.19 \times 10^3\ \text{J or } 1.19\ \text{kJ} \end{aligned}$$

The three-digit answer is 1.19 kJ.

8. Open Response

Improvements for product materials:

- use of a bomb calorimeter
- protect the flame from air currents (e.g.. surround it with a cylinder)
- include *m* and *c* for thermometer
- place in centre of calorimeter
- insulate calorimeter/ use stirrer
- t_i is as many °C below room temp. as t_f above room temp.
- do several trials
- calorimeter that conducts heat better than copper

9. –69.1

Hess's law is used to compute an enthalpy change for fermentation.

$$\Delta H^\circ_{\text{net}} = \sum nH^\circ_{\text{f (products)}} - \sum nH^\circ_{\text{f (reactants)}}$$

$$= [2\text{ mol} \times (-393.5\text{ kJ/mol}) + 2\text{ mol} \times (-277.6\text{ kJ/mol})]$$
$$[1\text{ mol} \times (-1\,273.1\text{ kJ/mol})] = -69.1\text{ kJ}$$

$$H^\circ_{\text{fermentation }C_6H_{12}O_{6(s)}} = \frac{\Delta H^\circ_{\text{net}}}{n_{C_6H_{12}O_6}}$$
$$= -69.1\text{ kJ/mol}$$

Or, since only one mole of glucose is fermented, the molar enthalpy of fermentation is −69.1 kJ/mol.

10. Open Response

$\Delta H_1 = [-1\,914.0\text{ kJ} + 2(-241.8\text{ kJ})]$
$-[-1\,129.7\text{ kJ} + 4(-271.1\text{ kJ})]$
$\Delta H_1 = -183.5\text{ kJ}$ (-271.5 kJ for water liquid)
$\Delta H_2 = -2\,112.9\text{ kJ} - (1\,914.0\text{ kJ})$
$\Delta H_2 = -198.9\text{ kJ}$
$\Delta H_{\text{net}} = -183.5\text{ kJ} + (-198.9\text{ kJ})$
$\Delta H_{\text{net}} = -382.4\text{ kJ}$

$$\frac{2.00 \times 10^6\text{g}}{352.03\text{ g/mol}} \times \frac{1}{1} \times 382.4\text{ kJ/mol} = 2.17\text{ GJ}$$

(2.19 GJ for U-235 with M = 349 g/mol)

11. Open Response

Variable/Change	Explanation
Temperature is decreased.	Lower fraction of effective collisions.

OR

Concentration of reactants was decreased.	Fewer collisions.

12. D

Unlike nuclear, wind, or solar power plants, a coal-burning plant relies on the combustion of fossil fuels. When these fuels burn, carbon gases are produced as waste. These carbon gases are greenhouse gases. They cause heat energy to become trapped in the atmosphere.

13. C

The reason a cold pack feels cold is that the reaction occurring within it absorbs kinetic energy from its surroundings. Formally speaking, when $NH_4NO_{3(s)}$ dissolves, more potential energy is **absorbed** in breaking ionic bonds within the salt than is **released** in forming bonds (ion dipole attractions) between water molecules and the NH_4^+ and NO_3^- ion, respectively.

Any reaction that cools its surroundings for this reason is an endothermic reaction. (Exothermic reactions warm their surroundings because more energy is released in forming bonds within the products than is absorbed to break bonds within the reactants).

14. B

The most likely role performed by vanadium oxide in the production of sulfuric acid is that of a catalyst. The best answer is alternative B.

Although it has no sulfur content, it does have a large number of oxygen atoms, making it unlikely to be a spectator ion or compound. Alternative A is therefore not the best answer.

A metal with five valence electrons would likely not bond with sulfate or sulfite ions because of the molar quantities of sulfur ions that would be required, and the number of bond formations would make such reactions very endothermic. C is therefore not a good answer.

Finally, vanadium oxide would likely have a high melting point, and when dealing with the production of acids, gases and liquids would probably be the states of matter in which most of the reactions would occur. The probability that it would act as a solvent is very low. D is therefore not a good answer.

15. D

B. Cellular respiration is strictly controlled so that an organism's moment-by-moment energy needs can be met. The rate increases with increased activity and decreases with decreased activity.

16. B

The enthalpy change of the hot chemical spray reaction is obtained by combining

i. one-half reversed reaction I
ii. reaction II reversed, and
iii. reaction III, unchanged, to give

$H_2O_{2(aq)} \rightarrow H_2O_{(l)} + \frac{1}{2}O_{2(g)}, \Delta H_I^\circ = -94.6\text{ kJ}$

$\frac{1}{2}O_{2(g)} + H_{2(g)} \rightarrow H_2O_{(l)}, \Delta H_{II}^\circ = -285.8\text{ kJ}$

$C_6H_4(OH)_{2(aq)} \rightarrow C_6H_4O_{2(aq)} + H_{2(g)}$
$\Delta H_{III} = +177.0\text{ kJ}$

Net:
$C_6H_4(OH)_{2(aq)} + H_2O_{2(aq)} \rightarrow C_6H_4O_{2(aq)} + 2H_2O_{(l)}$
$\Delta H^\circ_{\text{net}} = -203.4$

In using Hess's law in this way, the secret to a solution is to manipulate the given equations to give the compounds in the desired reaction in the correct proportions.
The elements usually remain in the correct proportions. Always remember that when a reaction is halved, doubled, tripled, etc., the enthalpy change is halved, doubled, tripled, etc., accordingly. If a reaction is reversed/flipped, the enthalpy change associated with it changes sign (from + to − and vice versa).

17. B

The formation of sulfur trioxide from sulfur dioxide can be written as

$SO_2 + \frac{1}{2}O_2 \rightarrow SO_3$

The first equation can be written as

$SO_2 \rightarrow S + O_2 \Delta H_1^{-} = +297\ \text{kJ}$

The required equation is obtained by adding the second equation to this equation.

$SO_2 + S + \frac{3}{2}O_2 \rightarrow SO_3 + S + O_2$

or, $SO_2 + \frac{1}{2}O_2 \rightarrow SO_3$

Thus, the enthalpy of formation of sulfur trioxide from sulfur dioxide can be calculated as

$\Delta H = \Delta H_2 + \Delta H_1^{-} = 297 - 396 = -99\ \text{kJ}$

Organic Chemistry

Organic Chemistry

Table of Correlations

Specific Expectation	Practice Questions	Self Test Questions
12.1.1 Understanding Basic Concepts		
12.1.1.1 *distinguish among the different classes of organic compounds, including alcohols, aldehydes, ketones, carboxylic acids, esters, ether, amines, and amides, by name and by structural formula*	1, 2	1
12.1.1.2 *describe some physical properties of the classes of organic compounds in terms of solubility in different solvents, molecular polarity, odour, and melting and boiling points*	6, 7	2
12.1.1.3 *describe different types of organic reactions, such as substitution, addition, elimination, oxidation, esterification, and hydrolysis*	8, 9	3, 4
12.1.1.4 *demonstrate an understanding of the processes of addition and condensation polymerization*	10, 11	5
12.1.1.5 *describe a variety of organic compounds present in living organisms, and explain their importance to those organisms*	14	6
12.1.3 Relating Science to Technology, Society, and the Environment		
12.1.3.1 *present informed opinions on the validity of the use of the terms organic, natural, and chemical in the promotion of consumer goods*	21, 22	12
12.1.3.2 *describe the variety and importance of organic compounds in our lives*	12, 13	13
12.1.3.3 *analyse the risks and benefits of the development and application of synthetic products*	23	
12.1.3.4 *provide examples of the use of organic chemistry to improve technical solutions to existing or newly identified health, safety, and environmental problems*	24	
12.1.2 Developing Skills of Inquiry and Communication		
12.1.2.1 *use appropriate scientific vocabulary to communicate ideas related to organic chemistry*	3	
12.1.2.2 *use the IUPAC system to name and write appropriate structures for the different classes of organic compounds, including alcohols, aldehydes, ketones, carboxylic acids, esters, ethers, amines, amides, and simple aromatic compounds*	4, 5	7
12.1.2.3 *build molecular models of a variety of aliphatic, cyclic, and aromatic organic compounds*	15, 16	8
12.1.2.4 *identify some nonsystematic names for organic compounds*	17	9
12.1.2.5 *predict and correctly name the products of organic reactions, including substitution, addition, elimination, esterification, hydrolysis, oxidation, and polymerization reactions*	18, 19	10
12.1.2.6 *carry out laboratory procedures to synthesize organic compounds*	20	11

12.1.1.1 *distinguish among the different classes of organic compounds, including alcohols, aldehydes, ketones, carboxylic acids, esters, ether, amines, and amides, by name and by structural formula*

12.1.2.1 *use appropriate scientific vocabulary to communicate ideas related to organic chemistry*

12.1.2.2 *use the IUPAC system to name and write appropriate structures for the different classes of organic compounds, including alcohols, aldehydes, ketones, carboxylic acids, esters, ethers, amines, amides, and simple aromatic compounds*

REVIEW OF ORGANIC CHEMISTRY

Organic chemistry is the study of chemical compounds that contain carbon atoms linked together in chains or rings. The simplest classes of organic compounds are hydrocarbons.
These compounds contain only carbon and hydrogen atoms. Hydrocarbons can be grouped into families. Alkanes, alkenes, alkynes, and cycloalkanes are all examples of these families.

FUNCTIONAL GROUPS

A **functional group** is an atom or group of atoms in an organic molecule that characterizes the molecule and enables it to react in specific ways. In other words, functional groups determine the chemical and physical properties of the molecules.

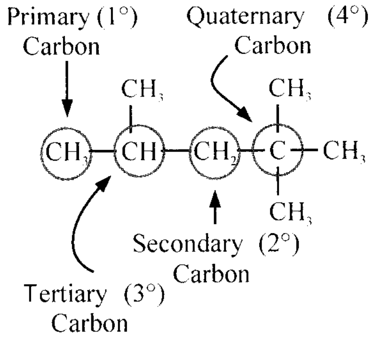
Functional groups are important because of the following reasons:

1. They are active sites for chemical reactions.
2. Each group is unique; therefore, you can classify organic compounds into families based on them.
3. They serve as a basis for naming organic compounds.

CLASSIFICATION OF CARBON AND HYDROGEN ATOMS

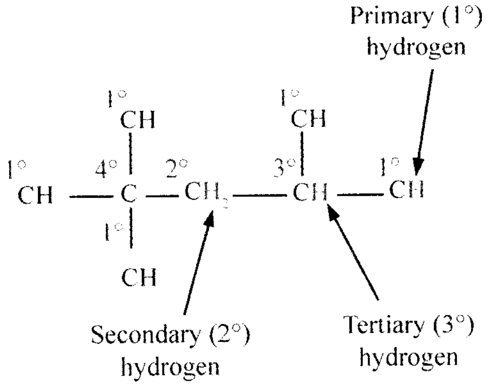
Carbon atoms can be classified according to the number of other C atoms that are bonded to it.

- Primary carbon: the carbon atom is directly attached to only **one** other carbon atom.
- Secondary carbon: the carbon atom is directly attached to only **two** other carbon atoms.
- Tertiary carbon: the carbon atom is directly attached to only **three** other carbon atoms.
- Quaternary carbon: the carbon atom is directly attached to **four** other carbon atoms.

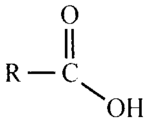
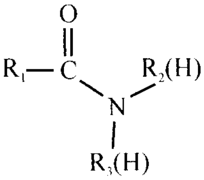
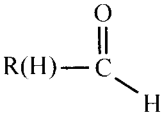
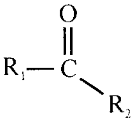


Similarly, a hydrogen atom is also classified as a primary, secondary, or tertiary hydrogen, depending on the type of C atom to which it is bonded.

- Primary hydrogen is bonded to a **primary** (1°) C
- Secondary hydrogen is bonded to a **secondary** (2°) C
- Tertiary hydrogen is bonded to a **tertiary** (3°) C
- There is no such thing as a quaternary hydrogen, because a quaternary (4°) C has no bonding sites left for an H to attach to.



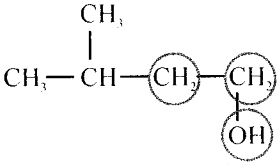
The families of organic compounds that you will study in this unit are summarized in the following table:

Class of Compound	Functional Group	
Alcohols	$R - OH$	Water Derivatives
Ethers	$R_1 - O - R_2$	
Amines	$R - NH_2$	
Carboxylic Acids	 $R-\overset{\overset{O}{\|}}{C}-OH$	Containing the C = H carbonyl group.
Amides	 $R_1-\overset{\overset{O}{\|}}{C}-\underset{\underset{R_3(H)}{\mid}}{N}-R_2(H)$	
Esters		
Aldehydes	 $R(H)-\overset{\overset{O}{\|}}{C}-H$	
Ketones	 $R_1-\overset{\overset{O}{\|}}{C}-R_2$	

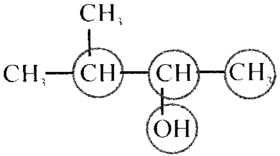
- R represents any alkyl group (any portion of a chain of a hydrocarbon).
- R_1, R_2, R_3 may or may not be the same alkyl group.
- R(H) means that the alkyl group, R, can be replaced with a hydrogen atom without seriously altering the structure of the functional group.

ALCOHOLS

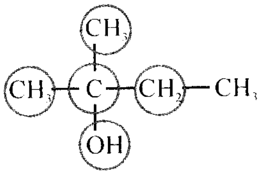
You can think of an alcohol (R – OH) as a water molecule in which one of the hydrogen atoms has been replaced by an alkyl group. Alcohols are characterized by having a hydroxyl (–OH) functional group attached to the C – H chain of an organic molecule. (From now on, reference to the C – H chain will be called the alkyl group.) There are 3 classes of alcohols: primary, secondary, and tertiary, depending on the carbon to which the hydroxyl group is directly attached.



Primary Alcohol
(OH attached to a primary C)



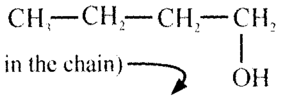
Secondary Alcohol
(OH attached to a secondary C)



Tertiary Alcohol
(OH attached to a tertiary C)

To name an alcohol, name the hydrocarbon as you normally would, following the rules for naming alkanes. The parent chain of the hydrocarbon should be numbered so that the carbon attached to the hydroxyl group has the lowest possible number. Drop the *-e* from the hydrocarbon name and replace it with *-ol*.

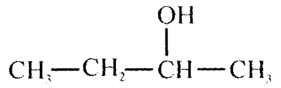
A number is added in front of the *-ol* suffix if the hydroxyl group is attached to a hydrocarbon with a parent chain larger than 2 carbons.



$$CH_3-CH_2-CH_2-\underset{\underset{OH}{|}}{CH_2}$$

(OH is on the first C in the chain)

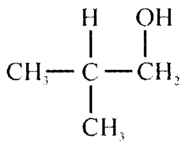
Butan-1-ol
(a 1° alcohol)
(formerly butyl alcohol)



$$CH_3-CH_2-\overset{\overset{OH}{|}}{CH}-CH_3$$

(OH is on the second C in the chain)

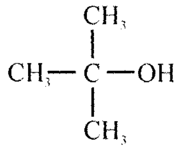
Butan-2-ol
(a 2° alcohol)
(formerly sec-butyl alcohol)



$$CH_3-\underset{\underset{CH_3}{|}}{\overset{\overset{H}{|}}{C}}-\overset{\overset{OH}{|}}{CH_2}$$

(OH is on the first C in the chain)

2-methylpropan-1-ol
(a primary alcohol)
(formerly isobutyl alcohol)
(methyl group is on the second C in the chain)



$$CH_3-\underset{\underset{CH_3}{|}}{\overset{\overset{CH_3}{|}}{C}}-OH$$

(OH is on the second C in the chain)

2-methylpropan-2-ol
(a tertiary alcohol)
(formerly tert-butyl alcohol)
(methyl group is on the second C in the chain)

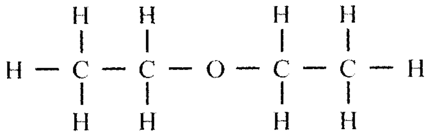
2-methylpropan-2-ol can also be named simply methylpropan-2-ol because the second carbon is the only place the methyl group can be added without changing the chain.

ETHERS

If you consider an alcohol (R – OH) to be a water molecule in which one of the hydrogen atoms has been replaced by an alkyl group, then you can think of an ether as a water molecule in which **both** of the hydrogen atoms are replaced by alkyl groups on either side. Ethers have the general formula R – O – R in which an oxygen atom is joined to two separate hydrocarbon groups.

Simple ethers are given common names in which the alkyl groups bonded to the oxygen are named in alphabetical order followed by the word *ether*.

$CH_3CH_2 - O - CH_2CH_3$ *Diethyl ether*



$$H-\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{C}}-\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{C}}-O-\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{C}}-\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{C}}-H$$

More complicated ethers require the use of oxy-derivatives of hydrocarbons

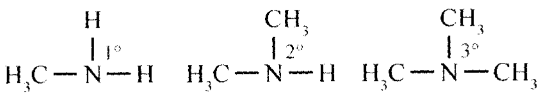
Alkyl Group	Name	Alkoxy Group	Name
CH_3-	Methyl	CH_3O-	Methoxy
CH_3CH_2-	Ethyl	CH_3CH_2O-	Ethoxy

$CH_3 - O - CH_2CH_3$	$CH_3CH_2 - O - CH_2CH_2CH_3$
Methoxyethane	Ethoxypropane
(Ethylmethyl ether)	(Ethylpropyl ether)

AMINES

Amines are organic compounds containing nitrogen. Amines are derivatives of ammonia. There are three classes of amines.

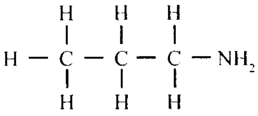
This classification is similar to the classification of alcohols since it is based on the number of carbon atoms (from alkyl groups) that are attached to the nitrogen atom.



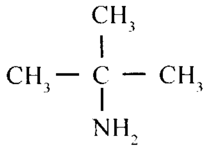
$$H_3C-\overset{\overset{H}{|\,1^\circ}}{N}-H \qquad H_3C-\overset{\overset{CH_3}{|\,2^\circ}}{N}-H \qquad H_3C-\overset{\overset{CH_3}{|\,3^\circ}}{N}-CH_3$$

1° Amine	**2° Amine**	**3° Amine**
Methylamine	Dimethylamine	Trimethylamine

To name an amine, name the hydrocarbon as you normally would, following the rules for naming alkanes. The parent chain of the hydrocarbon should be numbered so that the carbon attached to the nitrogen has the lowest possible number. Drop the *-e* from the alkane's name. Add the suffix *-amine* and a number specifying the carbon that carries the $-NH_2$ group.



Propan-1-amine



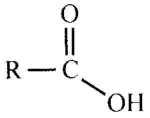
Propan-2-amine

The carbonyl group, C = O, is very common in biochemistry. It is the main component of the remaining functional groups that will be looked at.

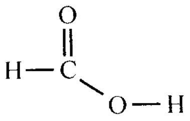
- Ketones and aldehydes (found in sugars/ carbohydrates)
- Carboxylic acids (found in amino acids)
- Esters (found in pheromones, scents and flavours in synthetic substances)
- Amides (found in proteins)

CARBOXYLIC ACIDS

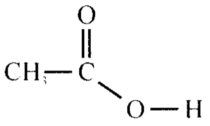
Carboxylic acids are compounds that contain a –COOH group.



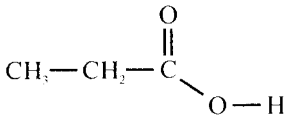
The name counts the total number of carbon atoms in the longest chain—including the one in the – COOH group. The parent chain of the hydrocarbon should be numbered so that the carbon attached to the –COOH has the lowest possible number. Drop the *-e* from the name of the alkane and add the suffix *–oic acid*. This functional group is always attached to the end of an alkyl group.



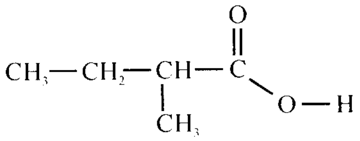
Methanoic Acid



Ethanoic Acid



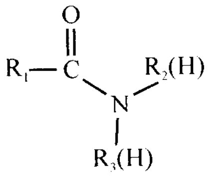
Propanoic Acid



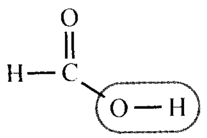
2-methylbutanoic Acid

AMIDES

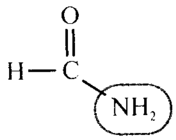
Amides are formed by the reaction of an amine with a carboxylic acid. In an amide, the –OH part of the carboxylic acid is replaced by an $-NH_2$ group when the amine and carboxylic acid groups react.



When naming an amide, name the compound as you normally would name the carboxylic acid, except replace the *-oic* acid ending with the suffix *-amide*.

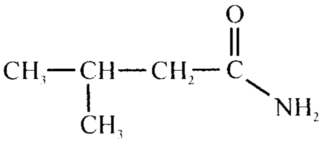


Methanoic Acid



Methanamide

The name counts the total number of carbon atoms in the longest chain—including the one in the $-CONH_2$ group. The parent chain of the hydrocarbon should be numbered so that the carbon attached to the $-CONH_2$ has the lowest possible number. This functional group is always attached to the end of an alkyl group.

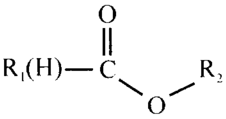


$CH_3-CH(CH_3)-CH_2-C(=O)-NH_2$

3-methylbutanamide

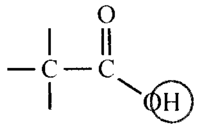
ESTERS

Esters are also derived from carboxylic acids, like amides. All esters have the general formula.



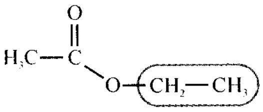
$R_1(H)-C(=O)-O-R_2$

A carboxylic acid contains the functional group $-COOH_2$. In an ester, the hydrogen from the $-COOH_2$ is most commonly replaced with another alkyl group.



$-C-C(=O)-OH$

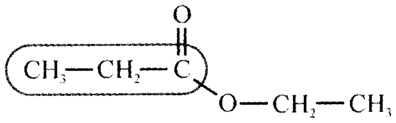
Ethanoic Acid



$H_3C-C(=O)-O-CH_2-CH_3$

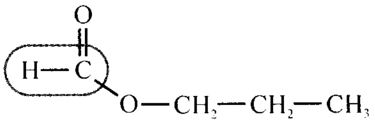
Ethyl Ethanoate

Naming an ester is quite tricky. First, name the carbon chain attached to the carbonyl group like you would any carboxylic acid. Notice that the acid is named by counting up the total number of carbon atoms in the chain—including the one in the $-COOH$ group. Remove the *-oic* acid suffix and replace it with an *-oate* ending. Next, determine the name of the alkyl group that replaced the hydrogen from the acid (methyl, ethyl, etc.) Notice that the ester is named the opposite way from the way the formula is written. The alkyl group name goes first and the *-oate* name is second, as shown below. (The *-oate* part of the structure in each molecule is circled.)



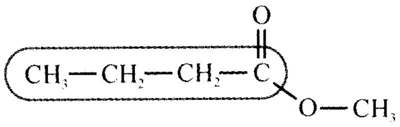
$CH_3-CH_2-C(=O)-O-CH_2-CH_3$

Ethyl Propanoate



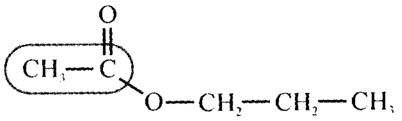
$H-C(=O)-O-CH_2-CH_2-CH_3$

Propyl Methanoate



$CH_3-CH_2-CH_2-C(=O)-O-CH_3$

Methyl Butanoate

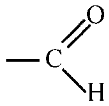


$CH_3-C(=O)-O-CH_2-CH_2-CH_3$

Propyl Ethanoate

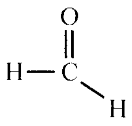
ALDEHYDES

All aldehydes contain the group .

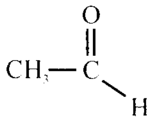


$-C(=O)H$

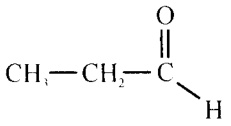
In aldehydes, the carbonyl group ($C = O$) has a hydrogen atom attached to it. This functional group is always attached to the end of an alkyl group.



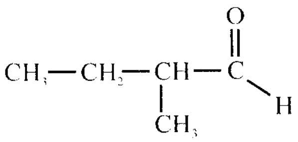
Methanal



Ethanal



Propanal

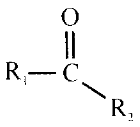


2-methylbutanal

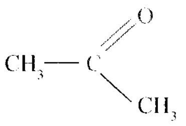
Note that the name "methanal" comes from "methyl aldehyde." It is very different from "methanol." When you are writing formulae for aldehydes, the functional group should always be written as $-CHO$, never as $-COH$, as this can easily be confused with an alcohol.
The name of an aldehyde counts the total number of carbon atoms in the longest chain—including the one in the $-CHO$ group. The parent chain of the hydrocarbon should be numbered so that the carbon attached to the $-CHO$ group has the lowest possible number.

KETONES

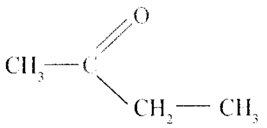
All ketones contain the group .



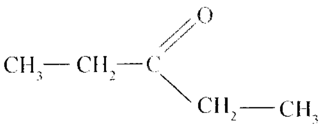
Ketones differ from aldehydes only by the fact that the carbonyl group has two hydrocarbon groups attached, rather than a hydrogen atom. In simple ketones, these hydrocarbon groups tend to be alkyl groups. The carbonyl group in a ketone cannot be on the end of a hydrocarbon, because that would make it an aldehyde.



Propanone



Butanone



Pentan-3-one

In naming a ketone, the parent chain of the hydrocarbon should be numbered so that the carbon attached to the carbonyl group has the lowest possible number. Drop the *-e* from the alkane's name. Add the suffix *-one* and a number specifying the carbon that carries the $C = O$ group.

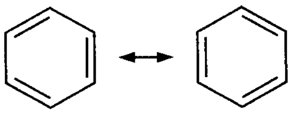
ALIPHATICS AND AROMATICS

Aliphatic compounds are composed of chains or branched chains of carbon atoms.

Aromatic compounds contain circular, ring-like structures of carbon atoms with a special bonding pattern, as seen in the molecular structure of benzene.

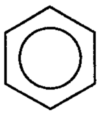
Benzene is an aromatic molecule that is formed from a carbon ring in a hexagonal structure. The six carbon atoms each have a single hydrogen atom bonded to them, and they are bonded to the carbon found on either side of them.

However, this only accounts for three of the four bonding electrons for each carbon atom. The extra electron from each of the six carbon atoms is shared within the hexagonal structure, forming neither single nor double bonds. The six electrons float about in a special area of probability in the shape of a ring. A way to visualize this is to imagine that each bond between two carbon atoms "vibrates" and alternates between a single and a double bond.



The symbolic representation of this arrangement is as follows:

Benzene
$C_6H_{6(l)}$

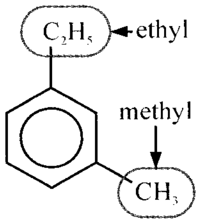


NAMING AROMATIC COMPOUNDS

If the parent chain of an aromatic hydrocarbon is a benzene ring, then the branches are attached to it, each branch replacing a hydrogen atom.
These branches are numbered so that the sum of the numbers is a minimum and in alphabetical order where necessary.

Example

Name the following compound.



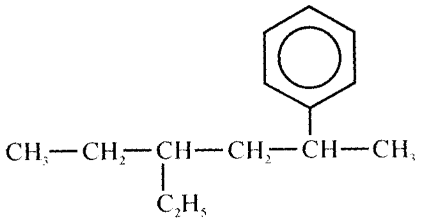
Solution

1. Alphabetically, ethyl precedes methyl.
2. In order to find the lowest numbers, assign the number 1 to the carbon bonded to the ethyl branch. Methyl would then be attached to the number 3 carbon.
3. The name of the compound is 1-ethyl-3-methylbenzene.

If a benzene structure is attached somewhere on a branched chain, it is usually treated as an addition to the chain, rather than as the parent chain. It is then called a **phenyl** group (not benzyl).

Example

Name the following compound.



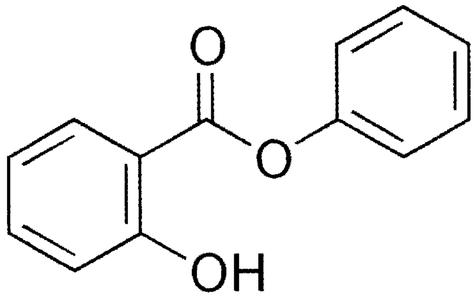
Solution

1. The phenyl group is attached to the second carbon on the chain.
2. There is an ethyl group attached to the fourth carbon.
3. The parent chain is heptane.
4. Alphabetically, ethyl precedes phenyl.
5. The name of the compound is 4-ethyl-2-phenylheptane.

Practice

Use the following information to answer the next question.

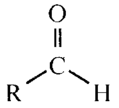
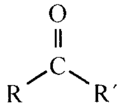
Salol was first introduced as a drug in 1886 by Polish chemist Marceli Nencki. It is used to bring down fever and as an antiseptic. The structure of Salol is shown below.



1. The two functional groups present in Salol are
   - **A.** a ketone and an ester
   - **B.** an amide and an amine
   - **C.** an ester and a hydroxyl
   - **D.** an amide and carboxylic acid

Numerical Response

Use the following information to answer the next question.

	Functional Group
1	R – O – R′
2	R – OH
3	
4	

2. The corresponding structures for ether, aldehyde, alcohol, and ketone are numbered __, __, __, and __, respectively.

3. Organic compounds that have the same functional group but different $-CH_2$ groups in their molecular formulae are known as
   - **A.** isomers
   - **B.** polymers
   - **C.** monomers
   - **D.** homologues

Open Response

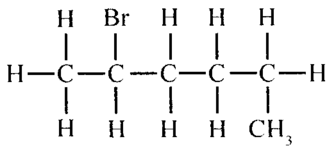
Use the following information to answer the next question.

Line structural formulae are the quickest way to provide all the necessary structural information about an organic compound. Carbon atoms are represented with one line for each bond to another atom. Hydrogens bonded to carbon are excluded, as their presence can be deduced from the carbon hybridization. It is important to draw all other atoms.

4. Draw the line structural formulae for 2-bromo-3-chlorohexane and pentane-1,3-diol. For each structure, note two instances of IUPAC nomenclature rules.

Open Response

Use the following information to answer the next question.



$$\mathrm{H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{Br}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{CH_3}{\overset{H}{C}}-H}$$

5. What is the IUPAC name of the given hydrocarbon?

12.1.1.2 *describe some physical properties of the classes of organic compounds in terms of solubility in different solvents, molecular polarity, odour, and melting and boiling points*

PHYSICAL PROPERTIES OF ORGANIC COMPOUNDS

The identity of the class of an organic compound can be determined qualitatively by observing the compound's physical properties. It is important to be familiar with the physical properties of the various groups of organic compounds. It is these properties that determine whether the compound is useful for a specific purpose in industry (i.e., high solubility makes for a good solvent). As well, finding patterns and trends in the physical properties of organic compounds provides you with evidence to explain what is happening at the molecular level (i.e., high boiling points = strong intermolecular bonds).

MELTING AND BOILING POINTS

Melting points (m.p.) and boiling points (b.p.) of all organic compounds are directly related to the intermolecular forces associated with that type of molecule. For example, when molecules are polar, they have dipole-dipole forces acting on them that weakly bond them together. These bonds increase the m.p. and b.p., since it now takes more energy to separate the molecules in the phase change.

Some polar molecules have hydrogen atoms attached to fluorine, oxygen, and nitrogen, which are highly electronegative elements (alcohols, for example –COOH). These types of molecules exhibit even higher m.p. and b.p. because hydrogen bonding occurs between molecules, increasing the energy required before a phase change can occur. All molecules have van der Waals dispersion forces acting upon them, polar or not. However, the larger a molecule becomes (the more carbon atoms it contains), the greater these dispersion forces become and the higher the amount of energy needed to break them. This, of course, increases the m.p. and b.p.

In every class of organic compound, you will find the following trends in m.p. and b.p.:

1. The larger the molecule, the higher the m.p. and b.p. (as a result of the increased length of molecule and therefore increased dispersion forces). Therefore, m.p. and b.p. always increase as the hydrocarbon chain gets longer.
2. Polar molecules have higher m.p. and b.p. than non-polar molecules.
3. Polar molecules containing hydrogen bonds have higher m.p. and b.p. than all other compounds.

SOLUBILITY

Solubility of organic compounds is also based on the polarity of the molecules. Polar molecules are soluble in water because of the fact that water is polar and can dissolve or hydrate hydrocarbons that are polar (like dissolves like). All hydrocarbons with functional groups attached are polar, and therefore are soluble in water to some degree. Aliphatic hydrocarbons (alkanes, alkenes, etc.) are non-polar, and therefore insoluble in water. **The solubility of polar organic compounds tends to decrease as the size of the molecule increases.**

ODOUR

Organic compounds are very often characterized by their odour. Many have b.p. at or near room temperature, and therefore give off fumes that are toxic (i.e., cyclobenzene). Care should be taken not to inhale the odours of most organic solvents. Esters, however, are synthesized specifically to produce a specific odour, many used in industry. Below is a list of some common esters and the odour they produce.

Ester	Odour
Ethyl ethanoate	Nail polish remover
Isobutyl methanoate	Raspberry
Methyl salicylate	Wintergreen
Methyl pentanoate	Flowers
Linalyl ethanoate	Lavender

Practice

Use the following information to answer the next question.

Nitrogen is less electronegative than oxygen. Therefore, amines form weaker hydrogen bonds as compared to alcohols and carboxylic acids. Carboxylic acids, on the other hand, have two oxygen molecules. Therefore, the extent of hydrogen bonding is higher in carboxylic acids as compared to amines and alcohols.

6. Which of the following compounds has the **lowest** boiling point?

A. C_3H_8

B. C_2H_5OH

C. HCOOH

D. $C_2H_5NH_2$

Open Response

7. Describe the factors that affect the solubility of gases, liquids, and solids in liquid solutions.

12.1.1.3 *describe different types of organic reactions, such as substitution, addition, elimination, oxidation, esterification, and hydrolysis*

SUBSTITUTION REACTIONS

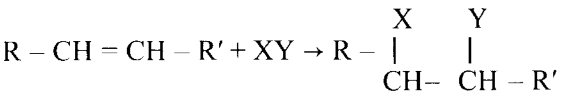
When a player in a baseball game or soccer game is replaced by another player, it is referred to as a substitution. Similarly, a substitution reaction is a reaction in which an atom or a functional group is replaced by a different atom or by a different functional group. This type of reaction can be represented as follows:

$$R - X + AY \rightarrow R - Y + AX$$

This is like a double displacement reaction that you studied in grade 10 and grade 11. This type of reaction is common in alkanes and in other saturated compounds, as well as in aromatic hydrocarbons.

ADDITION REACTIONS

Unlike saturated compounds, if you need to add an atom or a functional group to an unsaturated compound, you do not have to remove one of the existing atoms or functional groups as you do with a saturated compound. In an addition reaction, one of the bonds in the multiple bond breaks and two extra atoms or functional groups are added.
The atoms or groups may be the same or they may be different. These reactions are similar to the synthesis reactions that you studied in previous years. Alkenes and alkynes undergo such reactions. They can be represented as follows:

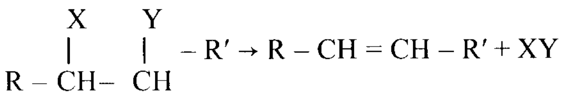


$$R-CH=CH-R'+XY \rightarrow R-\overset{\overset{X}{|}}{CH}-\overset{\overset{Y}{|}}{CH}-R'$$

If R and R′ are different, then two isomers can be formed.

ELIMINATION REACTIONS

An elimination reaction is the reverse of an addition reaction. In an addition reaction, two atoms or functional groups are removed (or eliminated) from a saturated molecule and a double bond is formed.



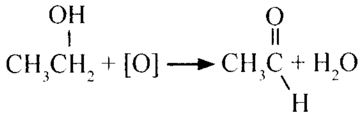
$$R-\overset{\overset{X}{|}}{CH}-\overset{\overset{Y}{|}}{CH}-R' \rightarrow R-CH=CH-R'+XY$$

This is similar to a decomposition reaction that you studied previously.

OXIDATION AND REDUCTION REACTIONS

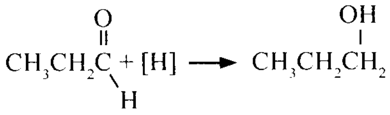
Over time, the definition of oxidation has been expanded to include any reaction in which electrons are lost. Reduction is defined as the gain of electrons. However, oxidation was originally defined as the addition of oxygen atoms or the removal of hydrogen atoms, and reduction was the opposite. In organic chemistry, an oxidation reaction is a reaction in which there is an increase in the number of bonds between carbon atoms and oxygen atoms or a decrease in the number of bonds between carbon atoms and hydrogen atoms. Reduction is the opposite.

A compound that increases the number of carbon-oxygen bonds or decreases the number of carbon-hydrogen bonds in another molecule is called an **oxidizing agent**. A **reducing agent** does the opposite. You represent oxidizing agents in a chemical equation as [O] and a reducing agent is represented as [H]. Oxidation is a common reaction of alcohols.



$$CH_3\overset{\overset{OH}{|}}{CH_2} + [O] \longrightarrow CH_3\underset{\setminus H}{\overset{\overset{O}{\|}}{C}} + H_2O$$

The alcohol is said to be **oxidized**.

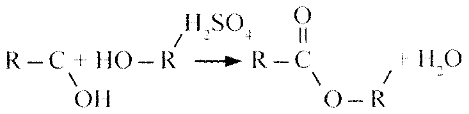


$$CH_3CH_2\underset{\setminus H}{\overset{\overset{O}{\|}}{C}} + [H] \longrightarrow CH_3CH_2\overset{\overset{OH}{|}}{CH_2}$$

The aldehyde is said to be **reduced**.

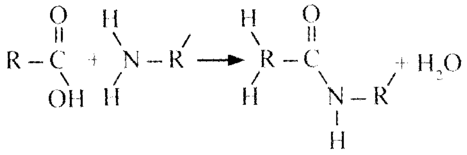
CONDENSATION REACTIONS

A condensation reaction is a reaction in which two organic molecules join together to form a larger molecule. A small molecule is also produced in a condensation reaction. This small molecule is commonly water. A typical condensation reaction is the reaction between an alcohol and a carboxylic acid. The main product of this reaction is an ester, and so this type of condensation reaction is called esterification. The reaction is catalyzed by H_2SO_4 or by H_3PO_4.



$$R-\overset{\displaystyle O}{\overset{\|}{C}}-OH + HO-R' \xrightarrow{H_2SO_4} R-\overset{\displaystyle O}{\overset{\|}{C}}-O-R' + H_2O$$

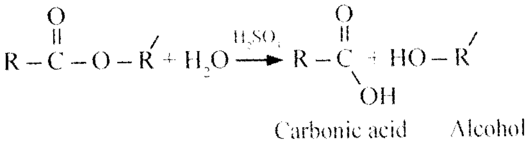
Another type of condensation reaction is the formation of an amide from a carboxylic acid and an amine. The reaction is similar to esterification, but the nitrogen atom of the amine takes the place of an oxygen atom.



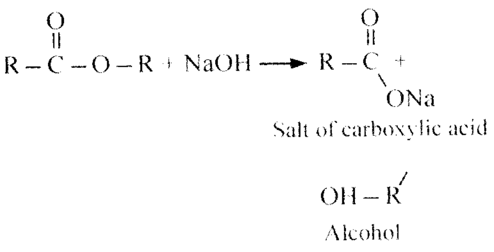
$$R-\overset{\displaystyle O}{\overset{\|}{C}}-OH + H_2N-R' \longrightarrow R-\overset{\displaystyle O}{\overset{\|}{C}}-NH-R' + H_2O$$

HYDROLYSIS

In an esterification reaction, an alcohol and a carboxylic acid combine to form an ester and water. When the reverse reaction occurs, i.e., water adds to an ester, causing it to split into an alcohol and a carboxylic acid, the reaction is called hydrolysis. The reaction can take place in the presence of an acid or a base.



$$R-\overset{\displaystyle O}{\overset{\|}{C}}-O-R' + H_2O \xrightarrow{H_2SO_4} \underset{\text{Carbonic acid}}{R-\overset{\displaystyle O}{\overset{\|}{C}}-OH} + \underset{\text{Alcohol}}{HO-R'}$$



$$R-\overset{\displaystyle O}{\overset{\|}{C}}-O-R + NaOH \longrightarrow \underset{\text{Salt of carboxylic acid}}{R-\overset{\displaystyle O}{\overset{\|}{C}}-ONa} + \underset{\text{Alcohol}}{OH-R'}$$

Practice

Use the following information to answer the next question.

When bromoethane is treated with aqueous sodium hydroxide, ethanol and sodium bromide are formed. The chemical equation involved in the reaction is as follows:

$$\underset{\text{Bromopropane}}{CH_3CH_2Br} + \underset{\text{Sodium Hydroxide}}{NaOH_{(aq)}} \rightarrow \underset{\text{Ethanol}}{CH_3CH_2OH} + \underset{\text{Sodium Bromide}}{NaBr}$$

8. The given reaction is an example of

A. addition reaction

B. elimination reaction

C. substitution reaction

D. esterification reaction

Open Response

Use the following information to answer the next question.

Esters are a class of organic compounds that often have a pleasant odour to them. Natural sources of esters include fruits and vegetables, flowers, and animal oils. Ester production is vital to the perfume industry and for developing artificial flavours. Ester syntheses can be accomplished by reacting carboxylic acids with alcohols. The alcohol loses a hydrogen atom and the acid loses the –OH portion of its carboxyl group. This reaction requires an acid catalyst.

9. Write a general formula for the esterification reaction, labelling each molecule type. (Use –R– to represent the alkyl groups.)

12.1.1.4 *demonstrate an understanding of the processes of addition and condensation polymerization*

12.1.3.2 *describe the variety and importance of organic compounds in our lives*

POLYMERS

All living organisms are primarily made up of things called polymers. Fingernails, hair, and most other parts of the body are polymers. A polymer is a **macromolecule** that is made up of small repeating units called monomers. The process that combines monomers to form polymers is called **polymerization**. Polymers can be divided into two types:

i. natural polymers

ii. synthetic polymers

Natural polymers are polymers that are not synthesized using artificial methods of construction, such as carbohydrates, proteins, etc.

Synthetic polymers are polymers that are prepared in industrial contexts to form things like plastics (e.g., Polythene, PVC, and Teflon) and fibres (e.g., polyester).

There are essentially two classes of polymers within the two types mentioned above:

- Homopolymers are composed of the same monomer.
- Copolymers are composed of two different monomers.

Homopolymers (Same monomers)	**Copolymers** (2 different monomers)
-A-A-A-A-A-A-A-	-A-B-A-B-A-B-A-
Linear	Linear
-A-A-A-A-A-A-A- | A-A-A	-A-B-A-B-A-B-A- | A-A-A
Branched	Branched
	-A-A-A-B-B-B-A-A-A-
	Blocked

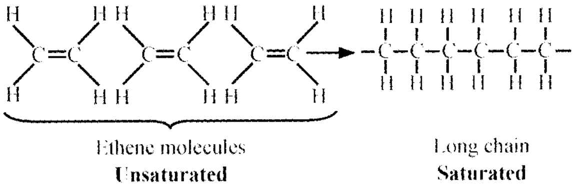
There are two main processes of polymerization: addition and condensation.

ADDITION POLYMERIZATION

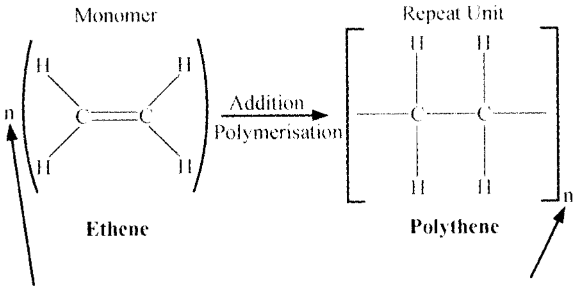
Under suitable conditions, it is possible to make monomers add together in large numbers to form chains without giving up any atoms. The process is called addition polymerization.

Example

When the monomer ethene goes through the process of addition polymerization, it forms polythene, a plastic material also known as polyethylene.



Notice that the double bonds between the carbon atoms break to "free up" a bonding electron and allow the monomers to join together. A "shorthand" equation is used to represent the polymerization process.

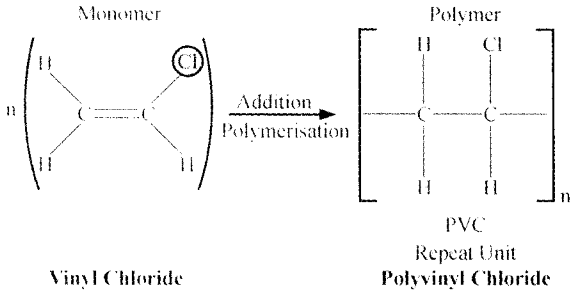


Here 'n' represents a very large number of ethene, C_2H_4 molecules joining together

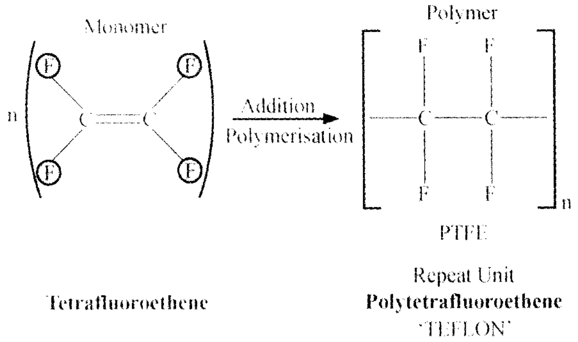
Here 'n' represents the same number of C_2H_4 units joined together

Ethene is used to make other polymer materials.

By replacing one hydrogen atom with a chlorine atom, **PVC** is produced.



Similarly, replacing all four hydrogen atoms with fluorine atoms produces **PTFE**.

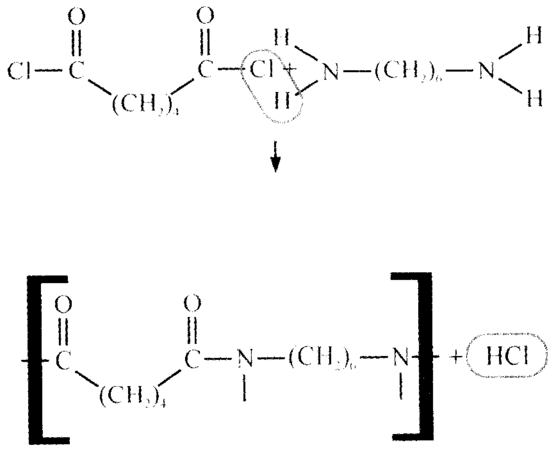


CONDENSATION POLYMERIZATION

Condensation polymerization only differs slightly from addition polymerization. In addition polymerization, double bonds are broken to allow bonding sites to become available. No particle or mass is lost. In condensation polymerization, monomers join together when certain atoms in the monomers get "kicked out" during the polymerization process. This expulsion of atoms frees up bonding sites and allows the monomers to bond together. As a result, condensation polymerization reactions give off byproducts, whereas addition polymerizations do not.
The byproducts, which most often tend to be HCl gas or water, are called *condensates*.
Typical condensation polymers are polyamides, polyesters, and certain polyurethanes.

Example

When one form of nylon is made from adipoyl chloride and hexamethylene diamine, the chlorine atoms from the adipoyl chloride, each along with one of the amine hydrogen atoms, are expelled in the form of HCl gas. Because there is less mass in the polymer than in the original monomers, you say that the polymer is *condensed* with regard to the monomers. The byproduct of this specific reaction, the HCl gas, is called a *condensate*.

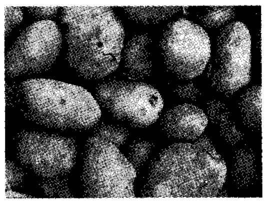


Plastics, such as the ones mentioned previously, are part of your everyday life; you use them constantly. Here are some other polymers than you come across every day, and as well as where you would see them.

Polymer	**Type of Polymer**	**Applications**
Polyethylene (PE)	Addition	Electrical wire insulation, flexible tubing, squeeze bottles
Polypropylene (PP)	Addition	Carpet fibres, ropes, liquid containers (cups, buckets, tanks), pipes
Styrofoam (Polystyrene)	Addition	Packaging foams, egg cartons, lighting panels, electrical appliance components
Polyvinyl chloride (PVC)	Addition	Bottles, hoses, pipes, valves, electrical wire insulation, toys, raincoats
Nylon (Polyamide)	Condensation	Fibres, clothing, tire cord, bearings
Kevlar (Polyamide)	Condensation	Bullet-proof vests, asbestos substitute, extreme sports equipment, composite aircraft construction.

Practice

Use the following information to answer the next question.



When certain fruits or potatoes are cut, the cut surface turns brown upon exposure to air. This is known as the enzymatic browning reaction. These dark brown chemical products are known collectively as melanins. Sodium sulfite is often used as a food additive because it prevents this enzymatic browning reaction.
Sodium sulfite is a member of a class of compounds called antioxidants.

10. Melanins are the product of polymerization of smaller molecules called ortho-quinones. In this regard, ortho-quinone could be considered

- **A.** a chain of repeating units
- **B.** a monomer
- **C.** a polymer
- **D.** an ester

Open Response

11. Compare and contrast addition polymerization and condensation polymerization reactions.

Use the following information to answer the next question.

A certain substance has the following properties:

- I. It is primarily composed of carbon.
- II. It burns in the presence of oxygen to produce huge amounts of energy.
- III. It is used to generate half of all the electricity in the United States of America.
- IV. It is used in industries for making steel, plastics, synthetic fibres, etc.

12. The substance is

A. coal
B. diesel
C. petrol
D. ethanol

Numerical Response

Use the following information to answer the next question.

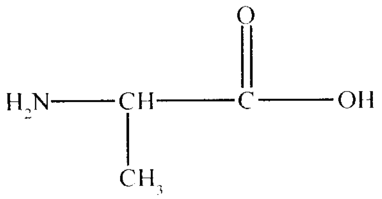
1. Automotive fuel
2. Shoe soles
3. Cardigan sweater
4. Baking soda
5. Table salt
6. Writing paper
7. Crystal glassware

13. Using the items given, list the three numbers that correspond to items **least** likely to contain organic compounds. ___, ___, ___.

12.1.1.5 *describe a variety of organic compounds present in living organisms, and explain their importance to those organisms*

BIOLOGICAL ORGANIC CHEMISTRY

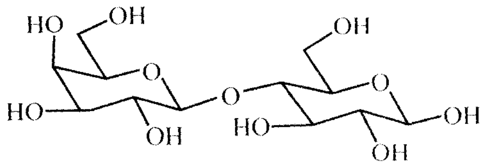
The four main nutrients in the human body are organic molecules that contain many of the functional groups discussed in this section. For instance, amino acids, the building blocks of protein in the human body, contain an amine functional group as well as a carboxylic acid functional group. For example, alanine is shown below:



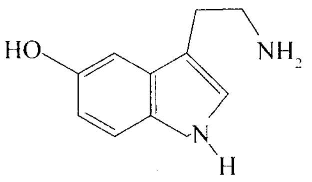
$$H_2N-\underset{CH_3}{\underset{|}{CH}}-\overset{O}{\overset{\|}{C}}-OH$$

Animal and vegetable fats are just big, complicated esters. When fats are synthesized for food production, they are just esters formed from the reaction between carboxylic acids and alcohols. If the acid used in the synthesis of the ester is saturated (no double bonds between any carbons), then the fat (ester) that forms will be a saturated fat. The acid used contains 1 double bond between 2 carbons; the resulting fat will be monounsaturated. If the acid used has more than one double bond between carbons, then the resulting fat will be a polyunsaturated fat. Remember that *unsaturated* means that hydrogen atoms can still be added.

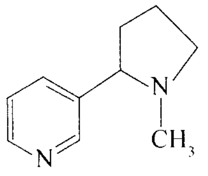
Carbohydrates are simple organic compounds that are aldehydes or ketones with many hydroxyl groups added, usually one on each carbon atom that is not part of the aldehyde or ketone functional group.



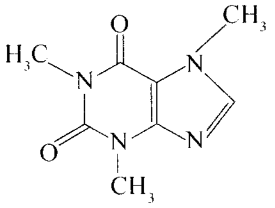
Organic compounds are everywhere in nature. Many nitrogen-based compounds are found naturally, particularly in plants and humans. They are called alkaloids.



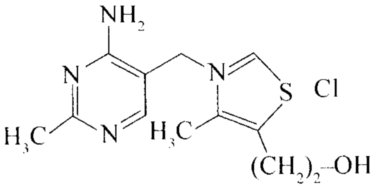
Serotonin
An important neurochemical



Nicotine (tobacco)



Caffeine



Thiamine
Vitamin B_1

Practice

Use the following information to answer the next question.

In living organisms, a certain widely prevalent organic compound is responsible for the growth of muscles. Body structures that are made from this compound include skin, hair, and nails. This compound is the sole provider of nitrogen to the body and is therefore a necessity in the diet.

14. The type of organic compound described is a

A. carbohydrate **B.** nucleic acid

C. protein **D.** fat

12.1.2.3 *build molecular models of a variety of aliphatic, cyclic, and aromatic organic compounds*

BUILDING MODELS OF ORGANIC COMPOUNDS

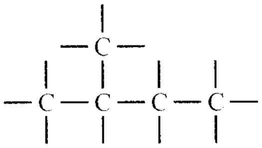
Organic molecules can be shown using three different types of diagrams: a complete structural diagram, a condensed structural diagram, and a line structural diagram.

TYPES OF DIAGRAMS

COMPLETE STRUCTURAL DIAGRAMS

Complete structural diagrams show all of the carbon atoms in an organic compound. Lines are drawn to indicate the bonds of each carbon atom, but the chemical symbols of any hydrogen atoms are not depicted. Any other chemical species present in the organic molecule are shown.

Example

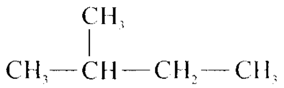


2-methylbutane

CONDENSED STRUCTURAL DIAGRAMS

In this type of diagram, all of the carbon-bonding groups are shown individually in molecular format, attached to one another in a line with branches representing attached chains of atoms.

Example



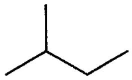
$$\overset{\displaystyle CH_3}{\overset{|}{}}\hspace{-1em}CH_3-\overset{\displaystyle CH_3\\|}{CH}-CH_2-CH_3$$

2-methylbutane

LINE STRUCTURAL DIAGRAM

Also known as a skeletal diagram, within this type of diagram, every junction between lines and the end point of each line indicates a carbon atom. There are enough hydrogen atoms bonded to each carbon to make the total number of bonds equal to four.

Example



2-methylbutane

ALIPHATIC STRUCTURES

Creating diagrams of organic compounds is important for several reasons, but the most salient one is to show the types of bonds present in a conformation. In other words, a molecular formula may show the number of atoms present in a molecule but not the arrangement of electrons, which will dictate the structure of the isomer of the molecule. Different bonds and arrangements of atoms will create different chemical properties that are manifested by the substance.

ALKANES

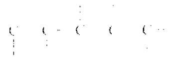
Molecular diagrams of single-bonded hydrocarbons use a single line to illustrate a bond between carbon atoms and other atoms, except in the case of line structural diagrams, which disregard hydrogen bonds.

Example

Given the molecular formula $C_5H_{12(l)}$, draw all three types of diagram for each of the three isomers.

Solution

Pentane



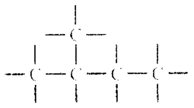
Complete structural diagram

$CH_3—CH_2—CH_2—CH_2—CH_3$

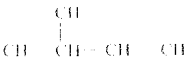
Condensed structural diagram

Line structural diagram

2-methylbutane

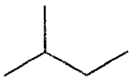


Complete structural diagram



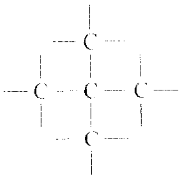
$CH_3\;\overset{\overset{\textstyle CH_3}{|}}{CH}—CH_2\;CH_3$

Condensed structural diagram

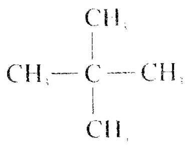


Line structural diagram

2,2-dimethylpropane

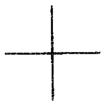


Complete structural diagram



$CH_3—\overset{\overset{\textstyle CH_3}{|}}{\underset{\underset{\textstyle CH_3}{|}}{C}}—CH_3$

Condensed structural diagram



Line structural diagram

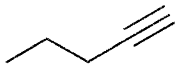
ALKENES AND ALKYNES

Organic molecules that have double and triple bonds within their structure are represented diagrammatically by double and triple lines between the symbolic depictions of the appropriate atoms. The existence of different bond types allows for more isomeric formations.

Example

Draw line structural diagrams for both isomers of pentyne $(C_5H_{8(l)})$.

Solution



pent-1-yne

pent-2-yne

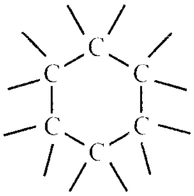
CYCLIC HYDROCARBONS

Organic compounds and functional groups that form closed, ring-like structures are generally represented diagrammatically as two-dimensional polygons. Double and triple bonds within those structures are represented with double and triple lines, respectively.

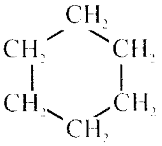
Example

Draw the structure of cyclohexane ($C_6H_{12(l)}$) using a complete structural diagram, a condensed structural diagram, and a line structural diagram.

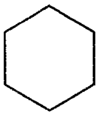
Solution



Complete structural diagram



Condensed structural diagram

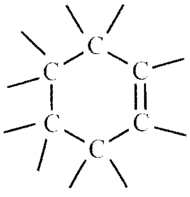


Line structural diagram

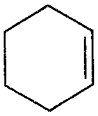
Example

Draw the structure of cyclohexene ($C_6H_{10(l)}$) using a complete structural diagram and a line structural diagram.

Solution



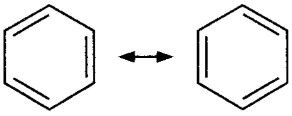
Complete structural diagram



Line structural diagram

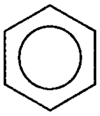
AROMATIC COMPOUNDS

Whether found as molecular benzene (C_6H_6) or as a phenyl functional group ($C_6H_5^+$), a hexagonal carbon ring with three alternating carbon bonds that form a resonance hybrid is depicted as a hexagon with a circle within it.



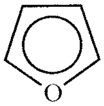
Resonance Hybrid

Benzene
$C_6H_{6(l)}$



Most aromatic compounds are based on a benzene ring structure, though there are some exceptions. A notable one is furan, the line structural diagram of which is shown below.

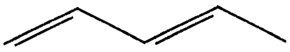
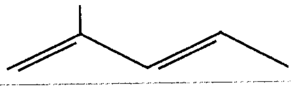
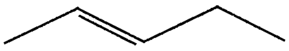
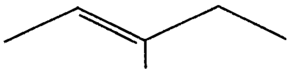
Furan
$C_4H_4O_{(l)}$



Practice

Use the following information to answer the next question.

The given table lists some alkenes along with their line diagrams.

Alkene (IUPAC name)	Bond Line Diagrams
Pent-1,3-diene	
2-methly-pent-1,3-diene	
Pent-2-ene	
2-methyl-pent-2-ene	

15. The line diagram of which alkene is listed **incorrectly** in the given table?

A. Pent-2-ene

B. Pent-1,3-diene

C. 2-methylpent-2-ene

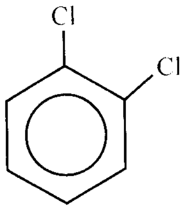
D. 2-methylpent-1,3-diene

Open Response

Use the following information to answer the next question.

When two functional groups are attached to a benzene ring, an alternate system to the IUPAC systematic naming system exists to identify the molecule.

The following molecule is systematically named 1,2-dichlorobenzene.



When the two functional groups are bonded to adjacent carbons, the compound can be named using the prefix *ortho*-. The alternate name for this compound is therefore ortho-dichlorobenzene. This form of notation can also be shortened to o-dichlorobenzene.

16. Draw and label the two structural isomers of o-dichlorobenzene.

12.1.2.4 *identify some nonsystematic names for organic compounds*

COMMON ORGANIC COMPOUND NAMES

You have learned how to name and draw structural formulae for organic compounds using the IUPAC system of nomenclature. However, many organic compounds have common names that differ from the IUPAC names. It is important to become familiar with these common names since they are used extensively in industrial and commercial sectors. Following is a chart that compares the IUPAC names with the common, non-systematic names of some compounds.

Systematic IUPAC Name	Non-Systematic Common Name	Uses
Ethanoic acid	Acetic Acid	Vinegar, industrial solvent
Propanone	Acetone	Industrial solvent, cleaning fluid, fuel additive
Propan-2-ol	Isopropyl alcohol	Electrical parts cleaner, preservative, industrial solvent
Methanal	Formaldehyde	Embalming, photography
Methanol	Methyl alcohol (wood alcohol)	Antifreeze, industrial solvent
Ethanol	Ethyl alcohol (grain alcohol)	Fuel, alcoholic beverages, antiseptic

Practice

17. Acetic acid is the common name of

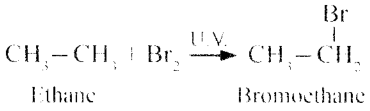
A. ethanoic acid **B.** butanoic acid

C. propanoic acid **D.** methanoic acid

12.1.2.5 *predict and correctly name the products of organic reactions, including substitution, addition, elimination, esterification, hydrolysis, oxidation, and polymerization reactions*

HALOGENATION OF ALKANES

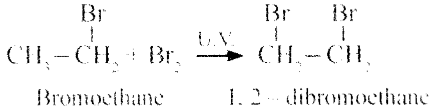
Alkanes can undergo substitution reactions with the halogens F_2, Cl_2, Br_2, and I_2. Except for the reaction with F_2, the process requires the addition of energy, either in the form of heat or ultraviolet radiation.



$CH_3-CH_3 + Br_2 \xrightarrow{U.V.} CH_3-\overset{Br}{CH_2}$

The product is named after the parent alkane with the appropriate prefix. To name the prefix, drop the *-ine* from the halogen name and replace it with *-o*. Hence, *bromine* becomes *bromo*,etc.

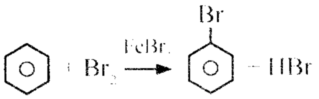
The substituted alkane can be reacted with more halogen, and a hydrogen atom bonded to another carbon atom is replaced.



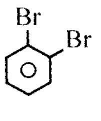
$CH_3-\overset{Br}{CH_2} + Br_2 \xrightarrow{U.V.} \overset{Br}{CH_2}-\overset{Br}{CH_2}$

When there are possibilities for isomers, the numbering system is used to indicate the position of the substitution. Use the prefixes *di-*, *tri-*, *tetra-*, etc., to show the number of substitutions.

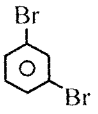
Benzene can undergo substitution with halogens in the presence of the catalyst $FeBr_3$.



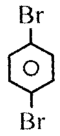
If the bromobenzene is reacted with more Br_2, then there are three possible products:



1, 2 dibromobenzene
ortho dibromobenzene
o - dibromobenzene



1, 3 dibromobenzene
meta dibromobenzene
m - dibromobenzene



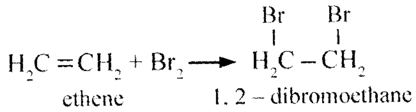
1, 4 dibromobenzene
para dibromobenzene
p - dibromobenzene

ADDITION REACTIONS

Alkenes undergo addition reactions with hydrogen, halogens, and hydrogen halides. When an alkene reacts with H_2 in the presence of a catalyst, it becomes saturated. The process is called **hydrogenation**.

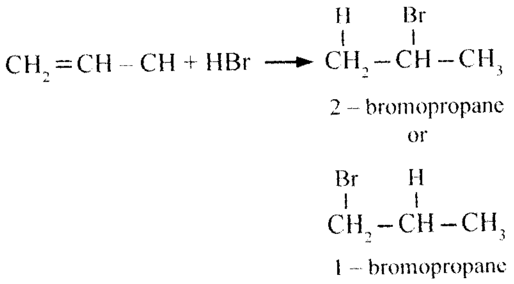
$$\underset{\text{alkene}}{H_2C=CH_2} + H_2 \xrightarrow{\text{catalyst}} \underset{\text{alkane}}{H_3C-CH_3}$$

A similar reaction occurs with halogens at room temperature without a catalyst.



$$\underset{\text{ethene}}{H_2C=CH_2} + Br_2 \longrightarrow \underset{\text{1, 2 - dibromoethane}}{H_2\overset{Br}{\overset{|}{C}}-\overset{Br}{\overset{|}{C}}H_2}$$

Alkenes react in the same way with hydrogen halides, like HBr. In this case, however, there are two possible products.



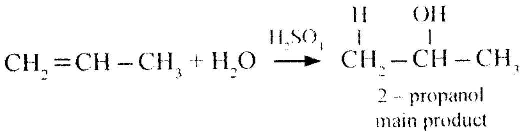
$$CH_2=CH-CH+HBr \longrightarrow \underset{\text{2 - bromopropane}}{\overset{H}{\overset{|}{C}}H_2-\overset{Br}{\overset{|}{C}}H-CH_3}$$

or

$$\underset{\text{1 - bromopropane}}{\overset{Br}{\overset{|}{C}}H_2-\overset{H}{\overset{|}{C}}H-CH_3}$$

2-bromopropane is the main product. There is a rule that tells you which is the main product in such cases. It is called **Markovnikov's rule** and it states that the carbon atom that is already bonded to the greater number of hydrogen atoms receives the hydrogen atom.

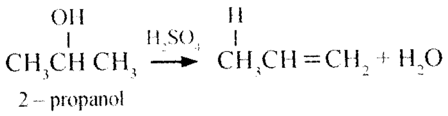
Similarly, alkenes can be hydrated using H_2SO_4 as a catalyst.



$$CH_2=CH-CH_3+H_2O \xrightarrow{H_2SO_4} \underset{\substack{\text{2 - propanol}\\ \text{main product}}}{\overset{H}{\overset{|}{C}}H_2-\overset{OH}{\overset{|}{C}}H-CH_3}$$

ELIMINATION REACTIONS

H_2SO_4 is also used as a catalyst to carry out the reverse of a hydrogenation reaction. Starting with an alcohol, you can eliminate water to produce an alkene.



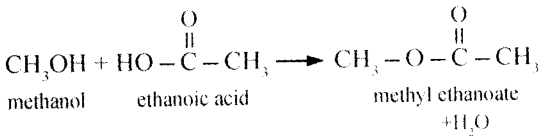
$$\underset{\text{2 - propanol}}{CH_3\overset{OH}{\overset{|}{C}}H\overset{H}{\overset{|}{C}}H_3} \xrightarrow{H_2SO_4} CH_3CH=CH_2+H_2O$$

The –OH group is removed from one C atom and an H atom is removed from its neighbouring C atom to form H_2O.

ESTERIFICATION

Esterification is a condensation reaction between an alcohol and a carboxylic acid. It is similar to but not the same as the neutralization reaction of NaOH with HNO_3.

NaOH + $\underline{H}NO_3$
base acid
sodium hydroxide nitric acid
→ $NaNO_3$ + H_2O
salt water
sodium nitrate

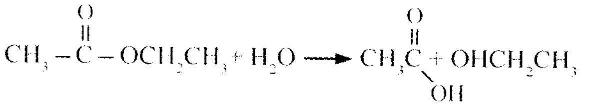


$$\underset{\text{methanol}}{CH_3OH} + \underset{\text{ethanoic acid}}{HO-\overset{O}{\overset{||}{C}}-CH_3} \longrightarrow \underset{\text{methyl ethanoate}}{CH_3-O-\overset{O}{\overset{||}{C}}-CH_3} + H_2O$$

To name the product, use the *alkyl* part of the alcohol name—for example, *methyl*, *ethyl*, etc. The second word of the name is named after the carboxylic acid, but drop the *-ic acid* part of the name and replace it with *-ate*, just as *nitric acid* becomes *nitrate*.

HYDROLYSIS

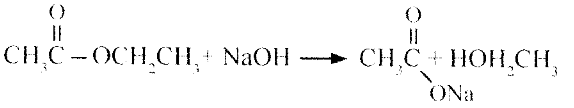
When the reverse reaction of esterification is carried out, the reaction is called **hydrolysis**.
Hydrolysis can be done in either acidic conditions or basic conditions.



$$\mathrm{CH_3-\overset{\overset{O}{\|}}{C}-OCH_2CH_3+H_2O\longrightarrow CH_3\overset{\overset{O}{\|}}{C}\underset{\diagdown OH}{}+OHCH_2CH_3}$$

ethyl ethanoate ethanoic acid ethanol

In base hydrolysis, the salt of the carboxylic acid is produced.



$$\mathrm{CH_3\overset{\overset{O}{\|}}{C}-OCH_2CH_3+NaOH\longrightarrow CH_3\overset{\overset{O}{\|}}{C}\underset{\diagdown ONa}{}+HOH_2CH_3}$$

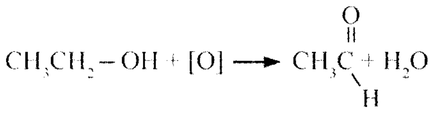
ethyl ethanoate sodium ethanoate ethanol

Base hydrolysis is
also called **saponification (soap-making).**

OXIDATION

Alcohols can be oxidized by oxidizing agents such as H_2O_2, $K_2Cr_2O_7$, and $KMnO_4$. In a chemical equation, an oxidizing agent is represented as [O]. Oxidation is defined as an increase in the number of bonds from a carbon atom to oxygen atoms or a decrease in the number of bonds to hydrogen atoms. A molecule that undergoes oxidation is said to be **oxidized**.

Primary alcohols are oxidized to aldehydes.

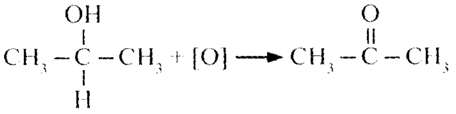


$$\mathrm{CH_3CH_2-OH+[O]\longrightarrow CH_3\overset{\overset{O}{\|}}{C}\underset{\diagdown H}{}+H_2O}$$

ethanol ethanal

Name the aldehyde by replacing the *-ol* in the alcohol name with *-al*.

Secondary alcohols are oxidized to ketones.

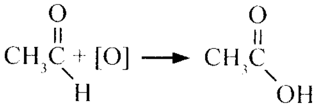


$$\mathrm{CH_3-\overset{\overset{OH}{|}}{\underset{\underset{H}{|}}{C}}-CH_3+[O]\longrightarrow CH_3-\overset{\overset{O}{\|}}{C}-CH_3}$$

2 – propanol propanone
+ H_2O

Name the ketone by dropping the *-e* from the name of the parent alkane and replacing it with *-one*.
If there are possible isomers, use the numbering system—e.g., 2-pentanone.

Aldehydes can be further oxidized to carboxylic acids.



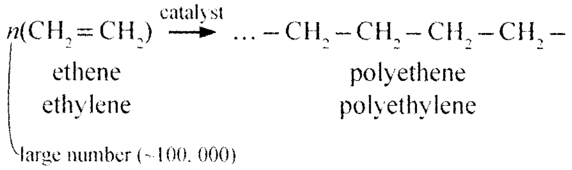
$$\mathrm{CH_3\overset{\overset{O}{\|}}{C}\underset{\diagdown H}{}+[O]\longrightarrow CH_3\overset{\overset{O}{\|}}{C}\underset{\diagdown OH}{}}$$

ethanol ethanoic acid

POLYMERIZATION REACTIONS

There are two types of polymerization reaction. Small double-bonded molecules can undergo addition to each other to form long chains called **addition polymers**. The single unit (small molecule) is called a **monomer**.

e.g.,



$$n(\mathrm{CH_2=CH_2})\xrightarrow{\text{catalyst}}\ldots-\mathrm{CH_2-CH_2-CH_2-CH_2}-$$

ethene polyethene
ethylene polyethylene
large number (~100, 000)

To name the polymer, put the prefix *poly-* in front of the monomer name.

The second type of polymer is called a **condensation polymer**. A molecule with two –OH groups can be reacted with a molecule with two carboxyl groups in multiple condensation reactions to form long chains of alternating alcohol and acid units. This type of polymer is called a **polyester**.

A similar type of reaction can be carried out using a molecule with two amino groups in place of a diol. The resulting polymer is called a **polyamide**.

Practice

18. What is the major product of the esterification reaction of propanoic acid with butanol?

A. Ethyl butanoate

B. Butyl butanoate

C. Butyl propanoate

D. Propyl propanoate

Open Response

19. Acetic acid and ethanol react through a process known as esterification. Predict and correctly name the products formed from this reaction.

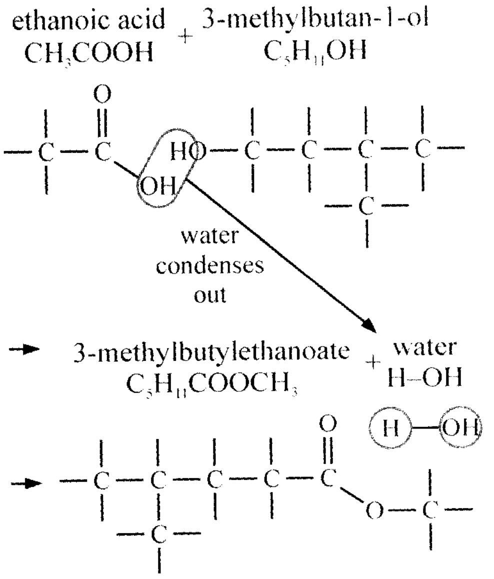
12.1.2.6 *carry out laboratory procedures to synthesize organic compounds*

SYNTHESIZING ORGANIC COMPOUNDS

ESTER SYNTHESIS

Esters are generally pleasant-smelling compounds. They are made by reacting alcohols with organic acids.

The ester formed when 3-methylbutan-1-ol is reacted with ethanoic acid is 3-methylbutylethanoate, which smells like banana oil.



Note: The process is not that simple: the reactants are usually heated for one hour, cooled, cold water is added, then sodium carbonate is used to remove excess acid. It is dried with magnesium sulfate, then finally distilled to purify it.

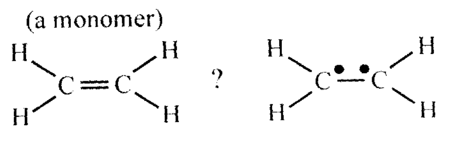
POLYMERIZATION

ADDITION POLYMERIZATION

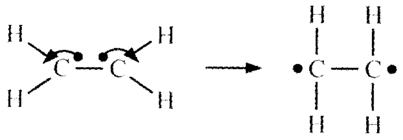
During the process of addition polymerization, molecules bond together to form a chain.

Many ethane molecules can be linked together by removing one pair of electrons from the double bond and using them to create bonds between the ethane molecules.

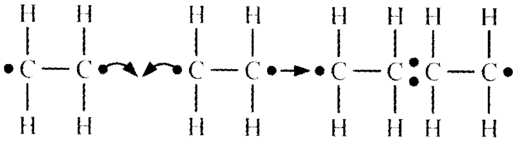
Each ethane molecule is a monomer in this example, and when they link together, the chain that is formed is known as a polymer.



Two electrons from the double bond have to "break free."



Next, the monomers can share those "free" unbonded electrons.



This results in a long chain, the polymer polyethene (also known as polyethylene). It retains its *-ene* name to show its origin, even though the product has no double bonds.

CONDENSATION POLYMERIZATION

The process of condensation polymerization involves connecting molecules together in a chain by removing one or two atoms from each end of each molecule, forming a chain and condensing a product, usually water.

There are many examples of natural condensation polymers, such as the following:

- Butter, which is a type of polyester
- Amino acid chains, which are polyamides
- Cellulose, which is a polysaccharide

There are also many common examples of synthetic condensation polymers:

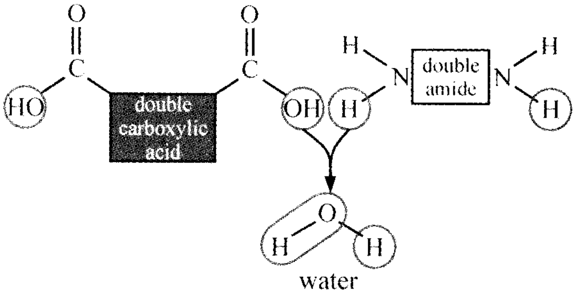
- Dacron, which is a polyester
- Nylon, which is a polyamide
- Ice-cream stabilizers, not all of which are natural —some are synthetic polysaccharides

In order to understand the making of nylon, it may be useful to visualize a wooden, miniature train track.

Wooden train track

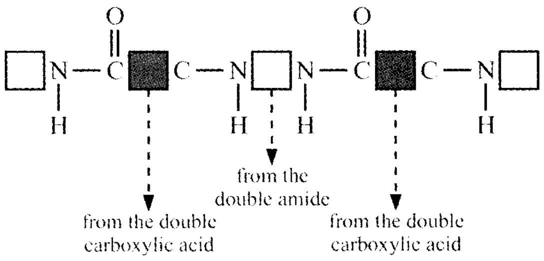
- Two types of monomer
- A is the same at both ends
- B is the same at both ends

Two kinds of monomer are used: one that has carboxyl (-COOH) groups at both ends (it is a double carboxylic acid) and the other that has amine groups at both ends (it is a diamine). If you ignore the carbon chains between the ends of each type of monomer, it might look something like this:



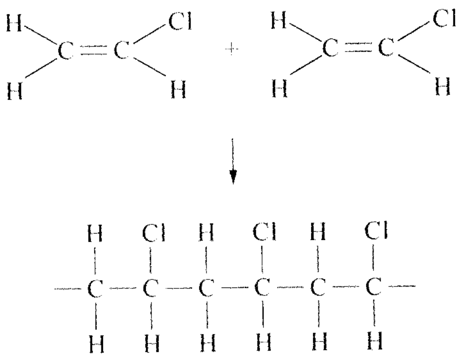
Formation of nylon

When the water is condensed out, the carbon and nitrogen atoms link together to form a chain like this.



Practice

Use the following information to answer the next question.



20. Which of the following terms **best** describes the given reaction?
   - **A.** Esterification
   - **B.** Elimination reaction
   - **C.** Addition polymerization
   - **D.** Condensation polymerization

12.1.3.1 *present informed opinions on the validity of the use of the terms organic, natural, and chemical in the promotion of consumer goods*

CHEMISTRY AND CONSUMERS

The word *chemical* is constantly used to describe harmful and dangerous substances that should be avoided at all costs. For example, many products you see on store shelves blatantly advertise the "absence" of chemicals with such phrases as "chemical-free", "all-natural", and "organic."

To most people, this is seen as a good thing. Unfortunately, there is a widespread misuse of terms such as these.

To start with, all things are made up of chemicals. The smell of a flower, the Rocky Mountains, the food you eat, the water you drink—all chemicals. Life in general is full of chemicals. Take the human body, for example. The minerals in bones, the hemoglobin in blood, and the neurotransmitters in the brain—all are chemicals. Without chemicals, life would cease to exist. It is therefore impossible for a substance to be chemical-free.

What are people thinking of when such bad images pop into their heads? It must be those nasty human-made chemicals, those synthetic chemicals found in food, air, and the environment that are so much more harmful to you than anything in nature. This unfortunately is another major misconception perpetuated every day. It is not the case that natural chemicals, i.e., those produced by plants and animals, are always good while human-made chemicals are always bad.

For example, the most toxic human-made chemical ever produced is called dioxin. This chemical has an LD_{50} of 3.0×10^{-2} mg/kg. (LD_{50} refers to the lethal dose for 50 % of the organisms exposed to it.) Even though this chemical is obviously quite potent, it is still a **million times** less toxic than the most toxic substance in the world, a naturally occurring substance called botulinum $\left(LD_{50} = 3.0 \times 10^{-8}\ \text{mg/kg}\right)$. This substance is so toxic that one teaspoon of it could kill a quarter of the world's population. (This is this same toxin that is used in the popular wrinkle-reducing procedure of Botox© injection.). In fact, five of the seven most deadly chemicals in the world are naturally occurring. These include tetanus toxin, diphtheria toxin, muscarine (found in mushrooms), and bufotoxin (found in the common European toad). Nature's toxins outrank those synthesized by chemists both in quantity and toxicity. Even some of the most deadly human-made toxins have benefits. Without the effectiveness of some human-made pesticides, the spread of malaria by mosquitoes would kill millions more people each year than it already does.

When it comes to food at least, people generally suppose that natural, organic food is better than the human-made, synthetic, genetically modified alternative. This is a misconception. Besides the fact that synthetic chemicals are, in fact, organic, if it were not for synthetic chemicals, food would not be safe to eat most of the time.

Without preservatives, food would not stay safe long enough for it to be shipped to supermarkets from other parts of the country or world. People should be more concerned with the diseases they could get from spoiled food than from the adverse effects of the preservative.

The availability of affordable, more stable, synthetic vitamins that are exactly the same as their natural counterparts is a benefit for humans. There is no scientific evidence to support the claim that organic food is better for you, although if organic farming means that pesticides are not used, there may be fewer traces of these pesticides on the food that reaches the market. If all food was grown purely organically with animal manure, you would need three times as many cows in the world to produce the fertilizer, which would cause many of its own problems.

Practice

Use the following information to answer the next question.

Isopentyl acetate ($C_7H_{14}O_2$) is the compound responsible for the scent of bananas. Bees also release a small quantity of this compound when they sting, signalling other bees to join the attack.

21. Isopentyl acetate is an organic compound because it

A. occurs naturally

B. contains oxygen

C. is produced by bees

D. is a molecular compound of carbon

Open Response

Use the following information to answer the next question.

Orlistat, also known as tetrahydrolipstatin, is a chemically-synthesized inhibitor of gastrointestinal lipases. It is used by people to lose weight because it reduces the amount of fat that can be absorbed by the digestive system. The synthesis is a carefully controlled five-step process designed to produce a very specific stereoisomer.

Orlistat has been marketed under two trade names by two large pharmaceutical corporations since 2003, and it has been scrutinized carefully by the drug administrations of a number of countries, with the United States only allowing over-the-counter sales in 2007.

Shortly thereafter, a number of companies changed one of the trade names by a single letter and began marketing an "herbal, organic" version of the drug.

22. Research the development and use of orlistat using the Internet, your library, and direct consumer research. In a brief essay, express the difference between the chemically synthesized version of the drug and the "herbal, organic" version. Which is more effective? Which is safer to use?

12.1.3.3 *analyse the risks and benefits of the development and application of synthetic products*

SYNTHETIC ORGANIC COMPOUNDS

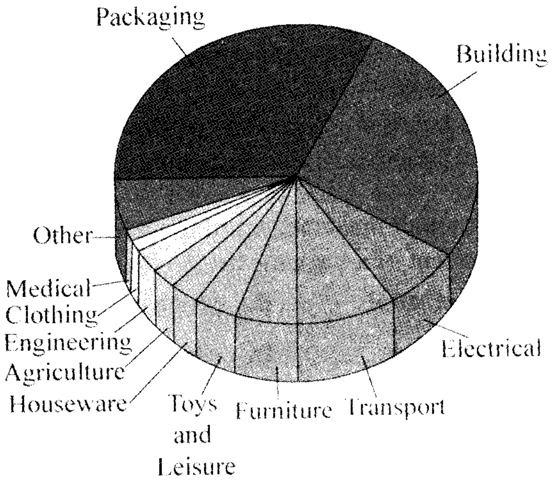
Organic materials (both natural and synthetic) play very important roles in people's everyday lives. Synthetic compounds are human-made organic molecules that are created in the laboratory. Some are replications of naturally occurring compounds (e.g., the vitamin C in a vitamin tablet is synthetic, yet it is the exact clone of the vitamin C molecule found naturally in citrus fruit—there is no difference whatsoever). Other synthetic compounds are created from scratch and are unique, not found in nature (e.g., the acetaminophen found in pain medication).

Polymers (such as polystyrene, nylons, and polythenes), solvents used in industrial processes, synthetic pesticides used to improve crop yields, and the myriad of pharmaceutical products used every day (aspartame in diet soft drinks, medicines) are all examples of synthetic organic compounds. These products have been researched and developed to make people's lives better. But with every benefit these compounds provide, there are always risks associated (e.g., aspartame allows you to have the taste of sugar without the caloric intake; however, scientists are still unsure of the long-term health risks that could be associated with it).

Plastics have become so popular that they now often replace traditional materials such as metals, wood, cardboard, and rubber in the manufacture of many consumer products. Their popularity has increased as a result of the many advantages they have over their traditional counterparts. For example:

- They can be easily coloured to look aesthetically pleasing.
- They can be moulded into complex shapes.
- They are strong yet light.
- They do not corrode when exposed to the atmosphere.
- They are cheap to produce.

From a social and economic perspective, plastics seem to be great. As shown in the given pie graph, plastics have a myriad of uses, from packaging of products to medical equipment.



So, what, if any, disadvantages does this wonder material have? From an environmental perspective, plastics are very difficult to get rid of. Most do not corrode, and they give off toxic fumes when burned. Unlike paper, most plastics are not biodegradable; they are not broken down by bacteria and stay structurally sound for hundreds of years. Because of this fact, most plastic waste ends up in landfill sites with limited capacity. Ironically, garbage is thrown away in plastic bags.

What are the solutions to this environmental nightmare? Scientists are working on new biodegradable plastics. Unfortunately, these cost more and are not economically viable at this time. Some countries are taxing the use of plastic bags or banning their use altogether. What about recycling? Even though recycling of plastics has increased hugely in the past decade, it has its disadvantages. Recycling uses a lot of energy, usually more than producing a new article from scratch. However, it does conserve the raw materials used originally.

Regardless of the type of synthetic material, there will always be risks and benefits associated with it. It is your obligation, as an informed citizen, to analyse these advantages and disadvantages and make informed decisions regarding their use.

Practice

Use the following information to answer the next question.

Aspartame, which contains phenylalanine, is a non-carbohydrate sweetener. It is a health hazard to people who are born with phenylketonuria, a genetic inability to process phenylalanine in the body.

23. In the small intestine, approximately 10 percent of aspartame breaks down into

A. methanol

B. propanol

C. ethanol

D. ethenol

12.1.3.4 *provide examples of the use of organic chemistry to improve technical solutions to existing or newly identified health, safety, and environmental problems*

Organic compounds play very important roles in people's personal lives. There are many examples of synthetic compounds that have very negative impacts on the environment. However, there are just as many examples of situations in which the use of organic chemistry has identified and provided solutions to environmental problems and massively reduced the damage done to the environment.

Example

In the 1980s, most aerosol sprays, fire extinguishers, and refrigerants were composed of haloalkanes (chemical compounds derived from an alkane by substituting one or more hydrogen atoms with halogen atoms). The most common of these organic halides were CFCs (chlorofluorocarbons).
After many years of use, it was found that CFCs had a direct link to the growing depletion of the ozone layer. Soon after this discovery, the international community passed the Montreal Protocol in 1987. This ground-breaking treaty set the guidelines for the phasing out of volatile organic compounds (VOCs), such as CFCs, worldwide over the years to come. Since its implementation, it has now become illegal to use CFCs in any form. Instead, they have been replaced with other haloalkanes, such as hydrofluorocarbons (HFCs). These compounds contain no chlorine. They are composed entirely of carbon, hydrogen, and fluorine. They have no known effects on the ozone layer. Only compounds containing chlorine and bromine are thought to harm the ozone layer.

Practice

24. Which of the following compounds would have the **least** environmental impact?

A. Chlorofluorocarbons

B. Hydrofluorocarbons

C. Bromofluorocarbons

D. Bromochlorocarbons

SOLUTIONS–ORGANIC CHEMISTRY

1. C	**6. A**	**11. OR**	**16. OR**	**21. D**
2. 1,3,2,4	**7. OR**	**12. A**	**17. A**	**22. OR**
3. D	**8. C**	**13. 4 5 7**	**18. C**	**23. A**
4. OR	**9. OR**	**14. C**	**19. OR**	**24. B**
5. OR	**10. B**	**15. C**	**20. C**	

1. C

The ester group, which has a general formula of RCOOR′, and the hydroxyl group (OH^-) are the two functional groups present in Salol.

2. 1,3,2,4

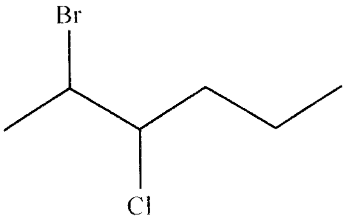
Ethers, aldehydes, alcohols, and ketones are characterized by their general formulae as follows:

Compound	General Formula	Functional Group
Ethers	R–O–R’	–O–
Aldehydes	R–CHO	–CH = O
Alcohols	R–OH	–OH
Ketones	R–CO–R	= O

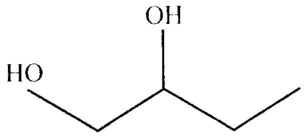
3. D

Homologues are compounds with the same functional group and elements. A homologue can be represented by a general formula, and each member differs from the adjacent members by one $-CH_2$ unit.

4. Open Response



2-bromo-3-chlorohexane



pentane-1,3-diol

Possible IUPAC rules for organic halides

1. Attached halogens are numbered in such a way to have the lowest numbers possible. It is not 5-bromo-4-chlorohexane.
2. The functional groups are listed in alphabetical order.
3. The alkane name represents the longest chain and is attached to the last halogen prefix as listed alphabetically.

Possible IUPAC rules for alcohols

1. The alkane name represents the longest carbon chain and is listed first.
2. Attached OH groups are numbered in such a way to have the lowest numbers possible. It is not pentane-3,5-diol.
3. The two alcohol functional groups are represented with the *di*- prefix.

5. Open Response

2-bromohexane

6. A

Since propane (C_3H_8) forms no hydrogen bonds, it has the lowest electronegativity. Since it is also non-polar, it will not exert any dipole-dipole attraction, and will only exhibit London forces. Therefore, its boiling point is the lowest.

7. Open Response

- "Like dissolves like" (polar solutes are more soluble in polar solvents—e.g., alcohol in water; non-polar solutes are more soluble on non-polar solvents —e.g., bike chain grease in white gas).
- Gases dissolve exothermically—this means that they tend to be more soluble at low temperatures and gases are more soluble under pressure—e.g., the $CO_{2(g)}$ in pop is more soluble when it is kept in the refrigerator with the lid on.
- Solids, which dissolve endothermically, tend to be more soluble at higher temperature—e.g., sugar is more soluble in hot coffee than it is in cold coffee.
- Liquids soluble in each other, in all possible ratios, are said to be miscible—e.g., methanol in water.
- Liquids that do not mix at all are said to be immiscible —e.g., oil and water.

8. C

Organic reactions are called substitution reactions when at least one molecular atom or functional group is substituted by another. Thus, the given reaction is an example of a substitution reaction. Here, the bromide (–Br) group in bromoethane is substituted or replaced by hydroxyl group (–OH) to form ethanol.

Addition reactions are those in which the reactants combine to form a single product.

Organic reactions are called elimination reactions when a specific part of a molecule separates to form a product.

In an esterification reaction, an alcohol reacts with a carboxylic acid to form an ester and water in the presence of a strong acid.

9. Open Response

$R - COOH + HOR' \rightarrow R - COOR' + H_2O$

10. B

A polymer is a large, usually organic, molecule built up from the repeated addition of small molecules, called monomers, to a growing chain either by condensation or radical chain-lengthening processes. Ortho-quinone must be the monomer from which melanins are made.

11. Open Response

Addition Polymerization

- The polymer is formed from many monomer subunits linked together.
- The monomers have double or triple bonds.
- The electrons in these multiple bonds rearrange before addition.
- The only product formed is the polymer.

Condensation Polymerization

- The polymer is formed from monomers that have at least two functional groups (one on each end of the molecule).
- The polymer is formed from the reaction of the functional groups of two different types of monomer molecules.
- In the process of forming the condensation polymer, some atoms are removed from each of the reacting functional groups to form a small molecule that is condensed out as a byproduct.

12. A

Coal is a fossil fuel that primarily contains carbon. The combustion of coal produces carbon dioxide and water. Huge amounts of energy are also released during the process. Half of the total electricity produced in the United States of America comes from the combustion of coal. Coal is also used in the manufacture of steel, plastics, synthetic fibres, etc.

Diesel and petrol are indeed obtained from crude oil. They produce huge amounts of energy when burnt. However, these are not used for the production of electricity and in the manufacture of steel, plastics, synthetic fibres, etc.

Ethanol is neither obtained by the fractional distillation of crude oil nor used in the generation of electricity.

13. 4 5 7

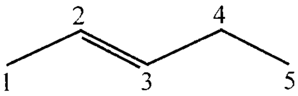
Petroleum, plastics, paper, and clothing usually contain large amounts of organic compounds. Table salt and baking soda are inorganic compounds, even though baking soda (sodium bicarbonate) contains carbon. The carbon in that particular molecule is found in a polyatomic anion, not in a covalent bonding arrangement. Crystal glassware is usually formed from silicate crystals, which are inorganic oxides of silicon mixed with metals such as lead.

14. C

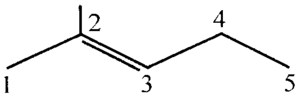
After water, protein is the organic compound that is the most prevalent in living organisms. Only proteins can provide nitrogen to the body.

15. C

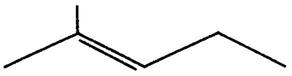
The alkene contains a parent chain with five carbon atoms, and the functional group $>C=C<$ is present at carbon-2 of the parent chain.



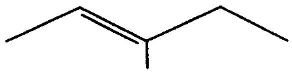
Methyl group is present at the carbon-2 of the parent chain.



Thus, 2-methylpent-2-ene has the given line diagram.



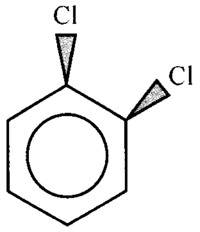
is the line diagram of 3-methylpent-2-ene.



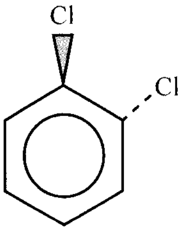
The bond line diagrams of pent-2-ene, pent-1,3-diene, and 2-methylpent-1,3-diene are listed correctly in the given table.

16. Open Response

Ortho-dichlorobenzene has two structural isomers: cis-ortho-chlorobenzene and trans-ortho-chlorobenzene. Their line structural diagrams are shown below.

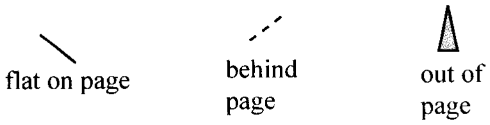


cis-ortho-chlorobenzene



trans-ortho-chlorobenzene

The three-dimensional nature of the bond formations are shown using the following key.



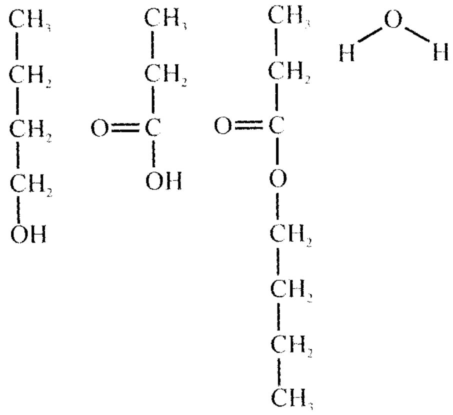
17. A

Acetic acid (CH_3COOH) is the common name of ethanoic acid.

18. C

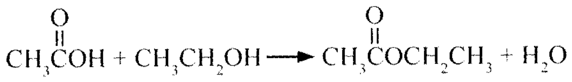
The reaction of propanoic acid with butanol is represented as follows:

butanol + propanoic acid → butyl propanoate + water



The major product of the reaction is butyl propanoate.

19. Open Response



$$CH_3\overset{O}{\overset{\|}{C}}OH + CH_3CH_2OH \longrightarrow CH_3\overset{O}{\overset{\|}{C}}OCH_2CH_3 + H_2O$$

The products are ethyl acetate and water.

20. C

This is an addition polymerization reaction of 1-chloroethylene. It is a version of polyethylene. Since no water is released as a product, it is not a condensation polymerization reaction.

21. D

Organic compounds are molecular compounds of carbon. Isopentyl acetate, the name of which labels it as an ester of acetic acid, is one such compound.

22. Open Response

Any response should include the difference in the standards of quality to which the manufacturers are held for the different products, the safety standards, marketing techniques, distribution, and biological impact of the drugs. There should also be some commentary on the different ingredients and reactants used to achieve the chemical effects the drug purports to achieve.

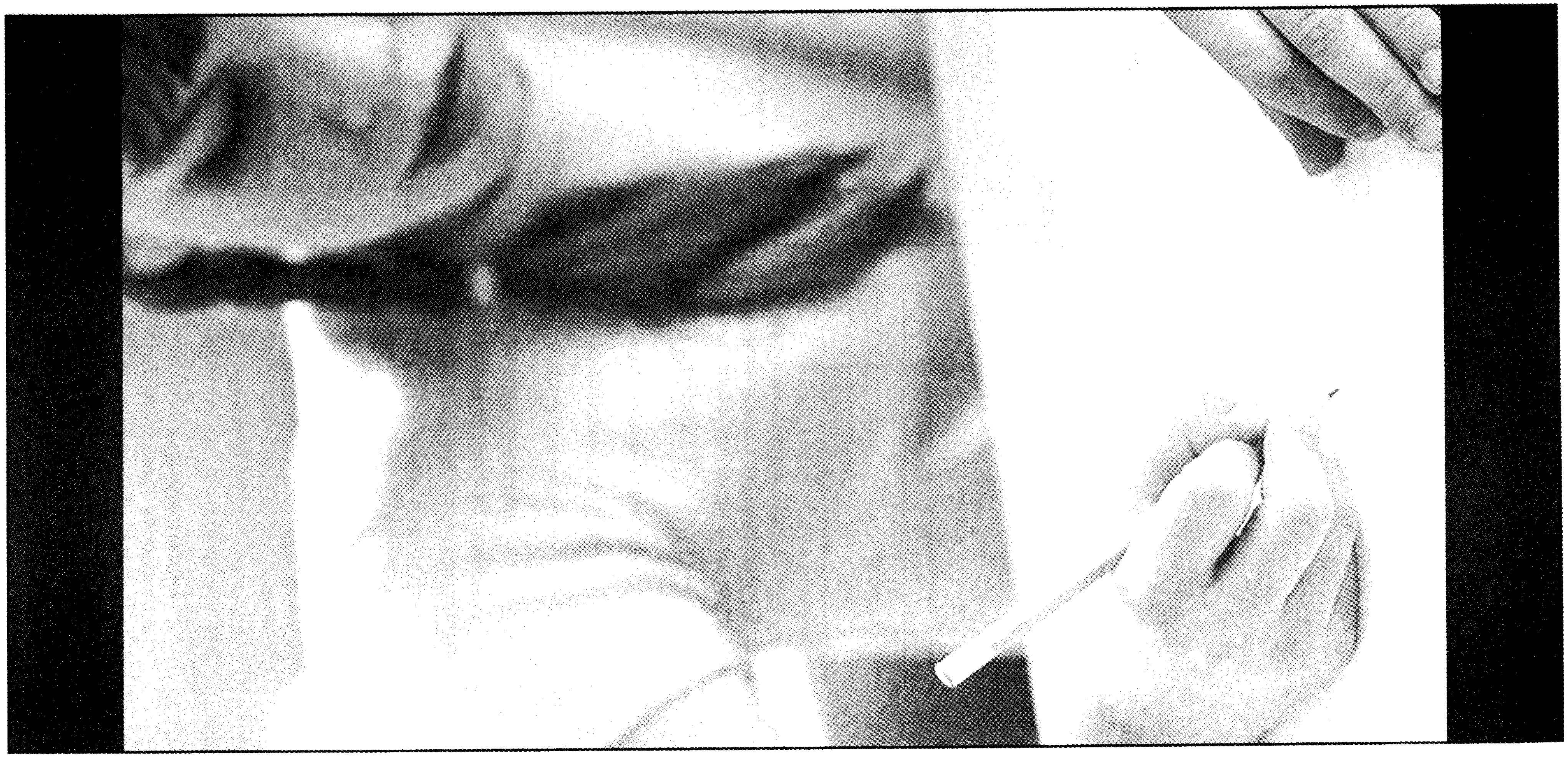
23. A

In the small intestine, approximately 10 percent of ingested aspartame breaks down into methanol through a hydrolytic process. Water molecules break apart the large aspartame molecule into smaller constituent compounds such as aspartic acid, phenylalanine, and methanol.

24. B

Organic halides that contain chlorine and bromine are thought to contribute to ozone layer depletion. Out of the four compounds listed, only hydrofluorocarbons do not contain either bromine or chlorine.

Self Test

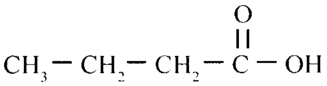


Use the following information to answer the next question.

Esters are the class of organic compounds that have pleasant smells. They are produced by the reaction of an acid and an alcohol in the presence of a strong acid. The reaction is called esterification. Esters are generally used in perfumes.

1. Which of the following organic compounds is an ester?

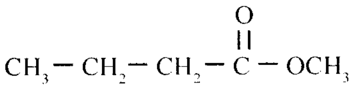
A.



$CH_3-CH_2-CH_2-\overset{\overset{O}{||}}{C}-OH$

B. $CH_3-CH_2-CH_2-CH_2-OH$

C.



$CH_3-CH_2-CH_2-\overset{\overset{O}{||}}{C}-OCH_3$

D. $CH_3-CH_2-CH_2-CH_2-Cl$

Use the following information to answer the next question.

Lower aldehydes and ketones containing up to four carbon atoms are soluble in water as a result of the presence of intermolecular hydrogen bonding with water molecules. With the increase in molecular mass, the polar carbonyl group is sterically hindered and its solubility decreases.

2. Which of the following carbonyl compounds is soluble in water?

A. Octanone
B. Dodecanal
C. Propanone
D. Benzaldehyde

Use the following information to answer the next question.

Ethanol is used as an alcoholic beverage and as an antifreeze in countries with colder climates. It is widely used in the chemical industry as a starting material for a number of other commercially important compounds. Ethene is prepared from ethanol by the following reaction:

$$CH_3CH_2OH \longrightarrow H_2O + H_2C=CH_2$$

3. The given reaction is an example of

A. an addition reaction
B. a substitution reaction
C. an elimination reaction
D. an esterification reaction

Open Response

4. Write an equation for the reaction between ethylene (ethene) and chlorine (Cl_2). Classify the reaction type.

Use the following information to answer the next question.

The polyethylene plastic used in film packaging and bottles is produced in a two-step reaction from petroleum.
Its manufacture can be represented as follows:

Step 1

$$C_2H_6 \xrightarrow{\text{catalyst}} CH_2 = CH_2 + H_2$$

Step 2

$$n CH_2 = CH_2 \xrightarrow{\text{catalyst}} \ldots CH_2 - CH_2 - CH_2 - CH_2 - \ldots$$

5. The conversion of ethene into a polymer in step 2 occurs by

- **A.** elimination
- **B.** substitution
- **C.** addition polymerization
- **D.** condensation polymerization

Open Response

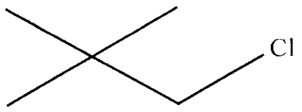
6. Is saturated or unsaturated fat more likely to undergo reactions with other molecules? Explain why.

Open Response

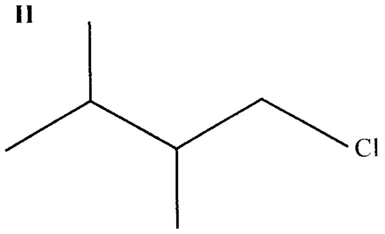
Use the following information to answer the next question.

Structures I and II are the line diagrams for two simple halogenated hydrocarbons.

I



II



7. What are the IUPAC names of the given compounds represented by structures I and II?

Open Response

8. Draw the full structural diagram of cyclohexane.

9. Which of the following compounds represents formic acid?

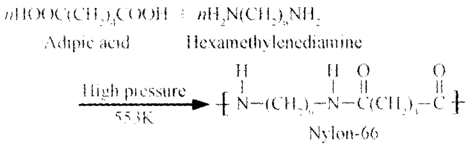
- **A.** HF
- **B.** $H_2S_2O_7$
- **C.** HCOOH
- **D.** CH_3COOH

Open Response

10. Two products are formed from the addition of HBr to propene. Predict and name the products.

Use the following information to answer the next question.

Nylon-66 is prepared by a polymerization of hexamethylenediamine with adipic acid. This reaction takes place under high pressure and temperature.



$$n\text{HOOC(CH}_2)_4\text{COOH} + n\text{H}_2\text{N(CH}_2)_6\text{NH}_2 \xrightarrow[553\text{K}]{\text{High pressure}} \left[\overset{\overset{\displaystyle\text{H}}{|}}{\text{N}}-(\text{CH}_2)_6-\overset{\overset{\displaystyle\text{H}}{|}}{\text{N}}-\overset{\overset{\displaystyle\text{O}}{\|}}{\text{C}}(\text{CH}_2)_4-\overset{\overset{\displaystyle\text{O}}{\|}}{\text{C}} \right]$$

11. Nylon-66 is formed by what type of polymerization reaction?

A. condensation **B.** addition

C. solution **D.** chain

12. Which of the following compounds is **not** a natural polymer?

A. Starch **B.** Protein

C. Cellulose **D.** Polyester

Open Response

Use the following information to answer the next question.

The labels on many of the products that people use today make a claim that the product is all natural. For example, an article of clothing may have a label that states that the article contains only natural fibres, or a food product may make the claim that it contains only natural flavourings.

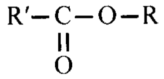
13. It is a commonly held belief that natural products are "good" and synthetic products are "bad." Use examples from at least three different industries (e.g., building materials, clothing, food, pharmaceuticals) to defend or criticize this belief.

SOLUTIONS

1. C	**5. C**	**9. C**	**13. OR**
2. C	**6. OR**	**10. OR**	
3. C	**7. OR**	**11. A**	
4. OR	**8. OR**	**12. D**	

1. C

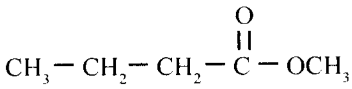
Esters are the compounds that contain a $-COOR$ group (R is the alkyl group). The general structure of an ester is



$$R'-\underset{\underset{O}{\|}}{C}-O-R$$

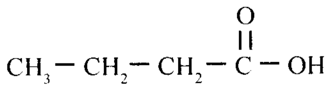
or R'COOR.

Therefore,



$$CH_3-CH_2-CH_2-\overset{\overset{O}{\|}}{C}-OCH_3$$

is an ester and its IUPAC name is methyl butanoate.



$$CH_3-CH_2-CH_2-\overset{\overset{O}{\|}}{C}-OH$$

is a carboxylic acid, as it contains a $-COOH$ group.

$CH_3-CH_2-CH_2-CH_2-OH$ is an alcohol, as it contains a hydroxyl group ($-OH$).

$CH_3-CH_2-CH_2-CH_2-Cl$ is a haloalkane, as it contains a chloro group.

2. C

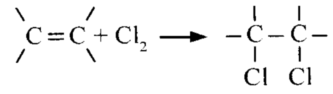
Since propanone has a low molecular mass, it will be soluble in water.

3. C

In the given reaction, a molecule of water is removed from the ethanol molecule, which results in the formation of a double bond. Therefore, the reaction is an elimination reaction.

In addition and substitution reactions, the number of molecular products would be equal to or less than the number of molecular reactants. Esterification reactions involve the formation of esters, which have a double-bonded oxygen within a carbon chain. Since both ethanol and ethene have only two carbons, an ester cannot be formed without adding more carbon atoms to the molecules involved.

4. Open Response



$$>C=C< + Cl_2 \longrightarrow -\underset{Cl}{\overset{|}{\underset{|}{C}}}-\underset{Cl}{\overset{|}{\underset{|}{C}}}-$$

This is an example of an addition reaction.

5. C

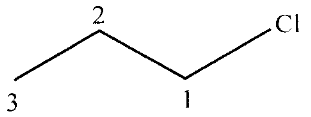
When many alkenes, like ethene in the given equation, are linked end-to-end to give a long chain alkane, the reaction is called addition polymerization.

6. Open Response

Unsaturated fat is the more reactive fat because it contains double bonds that are comparatively more reactive than single ones. The carbons in saturated fat are "saturated" with protons and are less likely to react further.

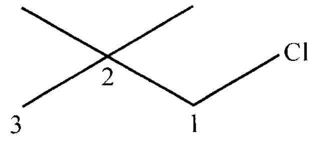
7. Open Response

Structure I: The parent chain contains three carbon atoms and the chlorine group is attached to the first carbon of the parent chain.



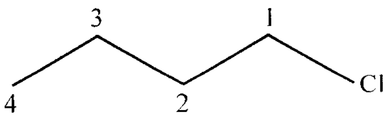
IUPAC name: 1-chloropropane

Two methyl groups are attached at carbon-2 of the parent chain.



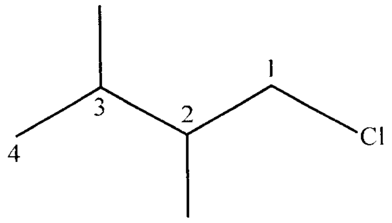
IUPAC name: 1-chloro-2, 2-dimethylpropane

Structure II: The parent chain contains four carbon atoms and the chlorine group is attached to the first carbon of the parent chain.



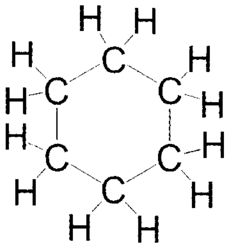
IUPAC name: 1-chlorobutane

The methyl groups are attached at carbon-2 and carbon-3 of the parent chain.



IUPAC name: 1-chloro- 2, 3-dimethylbutane

8. Open Response

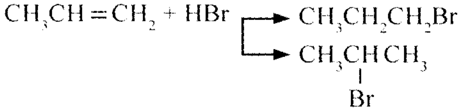


Cyclohexane

9. C

The formula of formic acid is HCOOH, and it is systematically named methanoic acid.

10. Open Response



$$CH_3CH{=}CH_2 + HBr \rightarrow \begin{cases} CH_3CH_2CH_2Br \\ CH_3\underset{\displaystyle Br}{\underset{|}{C}H}CH_3 \end{cases}$$

The two products are 1-bromopropane and 2-bromopropane.

11. A

Condensation polymerization involves the construction of a chain of repeating units with fewer constituent atoms than the reacting monomer species.

Hence, nylon-66 is formed by condensation polymerization of adipic acid and hexamethylenediamine. A is the correct answer.

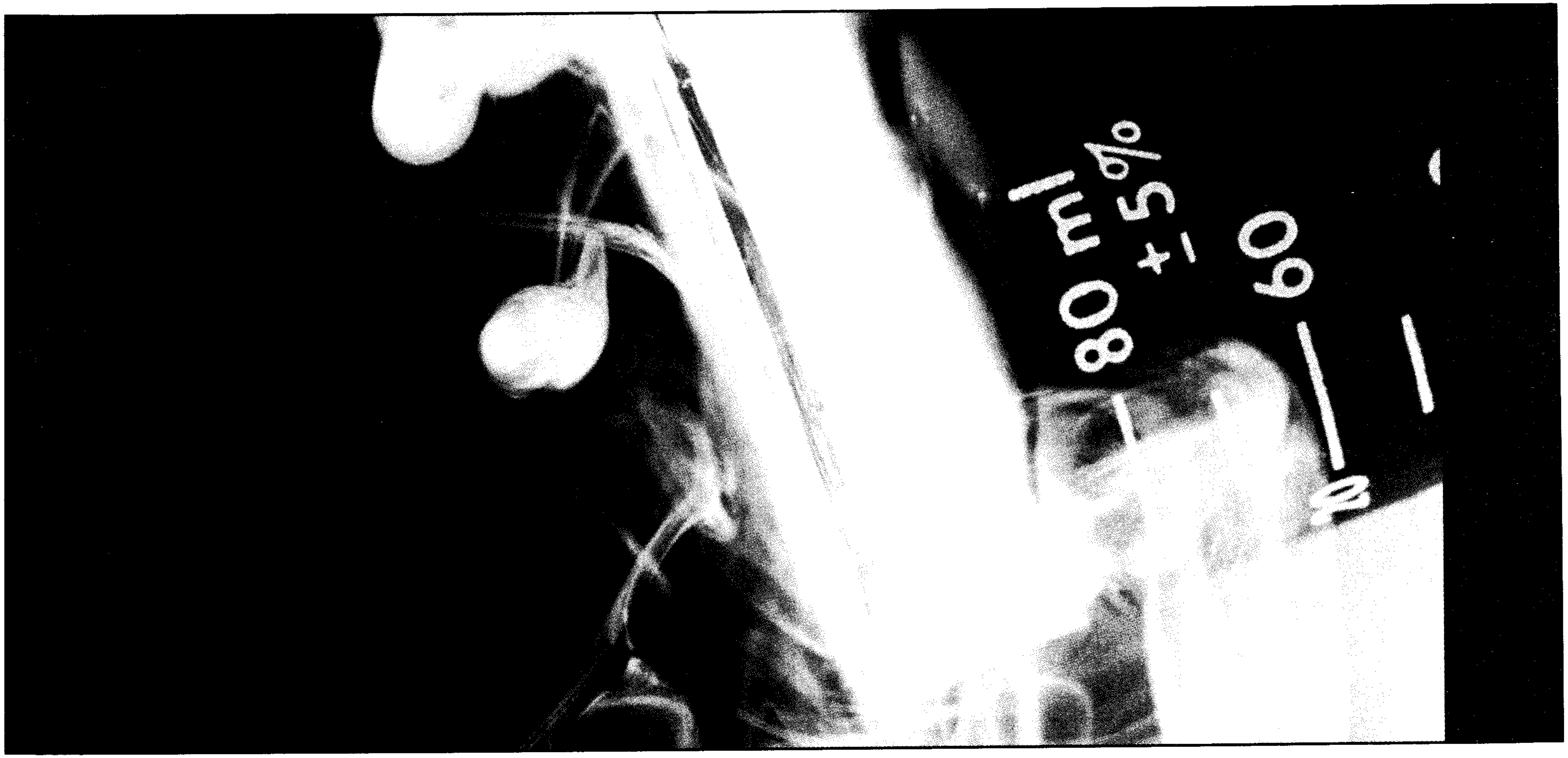
12. D

Polymers found in nature are called natural polymers, and polymers that are prepared in the laboratory are called synthetic or human-made polymers. Polyester is a synthetic polymer; the others are natural polymers.

13. Open Response

Any position advocated should demonstrate a knowledge of the terms involved and be substantiated with scientific evidence.

Systems and Equilibrium



Chemical Systems and Equilibrium

Table of Correlations

Specific Expectation	Practice Questions	Self Test Questions
12.3.1 Understanding Basic Concepts		
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12.3.3 Relating Science to Technology, Society, and the Environment		
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12.3.1.1 *illustrate the concept of dynamic equilibrium with reference to systems such as liquid-vapour equilibrium, weak electrolytes in solution, and chemical reactions*

DYNAMIC EQUILIBRIUM

Many reactions can easily progress in both the forward and reverse directions. These are called reversible reactions. Eventually, dynamic equilibrium is reached in reversible reactions. Equilibrium is the point in a reaction at which the rate of the forward reaction is equal to the rate of the reverse reaction. Molecules in the system are reacting continuously, but there is no net change in any properties of the system.

RATE FORWARD = RATE REVERSE

When writing a chemical equation, equilibrium is represented by a 2-way arrow ($\rightleftharpoons$). This symbolizes that the reaction is taking place in both forward and reverse directions simultaneously.

Example

$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

PROPERTIES OF EQUILIBRIUM

At equilibrium, there is no change in the macroscopic properties of the system. This includes no change in colour, no change in state, and no change in pressure. For gaseous and aqueous systems, there is also no change in the concentration of any substance. Equilibrium can only be attained in a closed system (no matter or energy can escape), but it can be reached from either the reactants or the products. Here are some examples of different types of systems that can reach equilibrium.

Example

System	Example
Liquid Vapour (PHASE)	$H_2O_{(l)} \rightleftharpoons H_2O_{(g)}$
Electrolyte	$Fe^{3+}_{(aq)} + SCN^{-}_{(aq)} \rightleftharpoons FeSCN^{2+}_{(aq)}$
Acid/Base	$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH^{+}_{4(aq)} + OH^{-}_{(aq)}$
Gaseous	$CO_{(g)} + Cl_{2(g)} \rightleftharpoons COCl_{2(g)}$
Dissolution (only if saturated)	$Li_2CO_{3(s)} \rightleftharpoons 2Li^{+}_{(aq)} + CO^{2-}_{3(aq)}$

Practice

Use the following information to answer the next question.

The burning of methane in a Bunsen burner to produce energy can be represented by the equation
$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$.

1. A student determined that the reaction represented by the given equation is **not** at equilibrium because

A. the system is open

B. a catalyst is not present

C. the temperature is constant

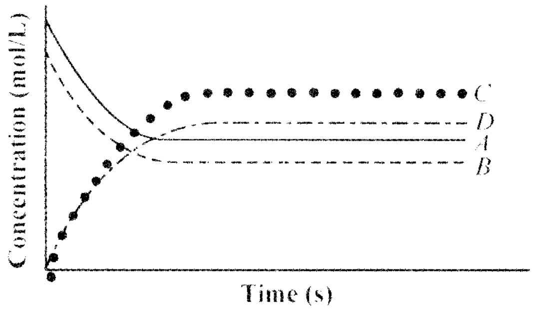
D. both reactants and products are gases

12.3.1.2 *demonstrate an understanding of the law of chemical equilibrium as it applies to the concentrations of the reactants and products at equilibrium*

THE EQUILIBRIUM CONSTANT

At equilibrium, the concentration of the reactants and the products does not change. On the graph below, equilibrium is observed where the concentrations of all species are no longer changing.

General Equilibrium Graph



This means that for the general equilibrium equation $aA + bB \rightleftharpoons cC + dD$, an equilibrium constant, K_c, can be written. If this expression is written for an equilibrium, it will always give the same constant value unless the temperature is changed. Notice that products, not reactants, are on the top of the K_c expression. It needs to be this way so that an equilibrium with a high concentration of products and low concentration of reactants will have a large value of K_c.

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

For the equilibrium $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$,

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

At a given temperature, the value of K_c remains constant despite stresses that are imposed on the equilibrium system. Because the concentration of products is on the top of the K_c ratio, the larger the value of K_c, the more the reaction favours products. In the reaction $aA + bB \rightleftharpoons cC + dD$ shown in the previous graph, K_c would be a large value, as there is a greater concentration of products than reactants at equilibrium.

GENERALIZATIONS

- If $K_c << 1$, the equilibrium favours the reactants.
- If $K_c >> 1$, the equilibrium favours the products.
- If $K_c \cong 1$, there are approximately equal concentrations of reactants and products at equilibrium.

Practice

2. For the given reaction $A + B \rightleftharpoons C + D$, what will happen when the reaction attains equilibrium?
   - **A.** The forward reaction as well as the backward reaction will stop.
   - **B.** The rate of the forward reaction will be greater than the rate of the backward reaction.
   - **C.** The rate of the backward reaction will be greater than the rate of the forward reaction.
   - **D.** The rate of the forward reaction will become equal to the rate of the backward reaction.

12.3.1.3 *demonstrate an understanding of how Le Châtelier's principle can predict the direction in which a system at equilibrium will shift when volume, pressure, concentration, or temperature is changed*

LE CHÂTELIER'S PRINCIPLE

Le Châtelier's Principle states that if a system at equilibrium is stressed by changing a reactant or product concentration, by changing the temperature, or by changing system volume and thus changing system pressure, the system will respond by shifting to oppose the imposed change.

If the equilibrium shifts such that more products are formed and more reactants are used up, the equilibrium is said to "shift right." If it shifts such that fewer products are formed and fewer reactants are used up, the equilibrium is said to "shift left." The following table summarizes the effects of equilibrium stresses.

Equilibrium Stress	Equilibrium Response	Affects the Value of K_c
Increase concentration of a reactant or Decrease concentration of a product	Shifts right	No
Decrease concentration of a reactant or Increase concentration of a product	Shifts left	No
Increase temperature of endothermic reaction or Decrease temperature of exothermic reaction	Shifts right	Yes, increases
Decrease temperature of endothermic reaction or Increase temperature of exothermic reaction	Shifts left	Yes, decreases

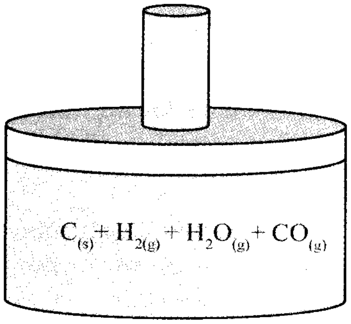
Equilibrium Stress	Equilibrium Response	Affects the Value of K_c
Increase pressure of a system by decreasing system volume	Shifts toward side with fewer moles of gases	No
Decrease pressure of a system by increasing system volume	Shifts toward side with more moles of gases	No

It is not necessary to memorize this table; an understanding of Le Châtelier's principle allows you to predict these effects. Simply picture the equilibrium fighting back whenever you do something to it.

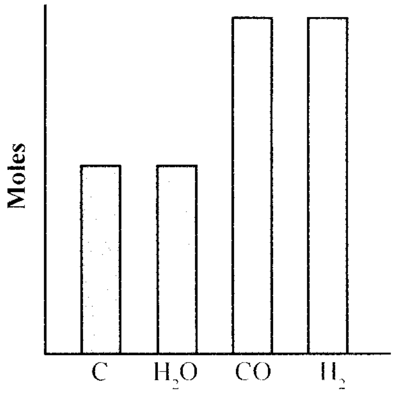
Practice

Use the following information to answer the next question.

The given figure shows a piston cylinder that contains 0.06 mol of carbon, water, carbon monoxide, and hydrogen each. The volume of the cylinder is 5.0 L.



The given graph represents the concentration of the various species at equilibrium.



The reaction reaches an equilibrium when heated to a temperature of 1 000 K.

$C_{(s)} + H_2O_{(g)} \rightleftharpoons CO_{(g)} + H_{2(g)}$

3. When the concentration of CO is decreased, what effect will this have on the equilibrium of the system?
   - **A.** The reaction will go to completion, turning all available reactants to products.
   - **B.** The concentration of hydrogen gas will increase.
   - **C.** The concentration of water vapour will increase.
   - **D.** The state of equilibrium will not be affected.

12.3.1.4 *identify, in qualitative terms, entropy changes associated with chemical and physical processes*

ENTROPY

A spontaneous reaction occurs without continuous outside assistance. Many people think that all exothermic reactions are spontaneous, but this is not true. Spontaneity does depend on enthalpy, ΔH, but also entropy, ΔS. Entropy, ΔS, is a measure of the randomness of disorder of a system with the units J / K. The more disordered a system is, the greater its entropy.

In general,

- if $\Delta S > 0$, there is more disorder and the reaction tends to be spontaneous.
- if $\Delta S < 0$, there is less disorder and the reaction tends not to be spontaneous.

Entropy increases in two ways:

- a change of state reflecting more disorder ($S_{\text{solid}} < S_{\text{liquid}} < S_{\text{gas}}$)
- few large molecules produce many small molecules

For example, in cellular respiration there is an increase in entropy, $\Delta S > 0$, because there are more small molecules on the product side and there is greater disorder in the state on the product side.

$C_6H_{12}O_{6(s)} + 6O_{2(g)} \rightleftharpoons 6CO_{2(g)} + 6H_2O_{(l)}$

REACTION SPONTANEITY

The standard free energy change (a.k.a. Gibbs free energy), ΔG, of a reaction must be negative for a reaction to be spontaneous, or $\Delta G < 0$.
The relationship between enthalpy and entropy to determine if a reaction is spontaneous or not is summarized as follows, where T represents temperature: $\Delta G = \Delta H - T\Delta S$.

Therefore, when determining if a reaction is spontaneous or not, it is important to know both the sign of ΔH and ΔS. The following table summarizes the use of this equation to determine spontaneity.

	Endothermic $\Delta H > 0$	**Exothermic** $\Delta H < 0$
Entropy increases $\Delta S > 0$	Spontaneous at high T	SPONTANEOUS
Entropy decreases $\Delta S < 0$	NOT SPONTANEOUS	Spontaneous at low T

Example

$PCl_{3(g)} + Cl_{2(g)} \rightleftharpoons PCl_{5(g)} + 156.5\ \text{kJ}$

$\Delta H < 0 \rightarrow$ Spontaneous

$\Delta S < 0 \rightarrow$ Not spontaneous

Therefore, the reaction $\Delta G < 0$ is spontaneous at low temperatures.

Practice

Open Response

4. For the following reaction, discuss two ways the entropy of the system changes as the reaction progresses toward the reactants.

   $C_6H_{12}O_{6(s)} + 6O_{2(g)} \rightleftharpoons 6CO_{2(g)} + H_2O_{(l)}$

12.3.1.5 *describe the tendency of reactions to achieve minimum energy and maximum entropy*

MINIMUM ENERGY

You are familiar with everyday examples of systems moving from higher energy to lower energy.

It takes energy to lift a ball from the floor to chest height, but if you let go, it returns to a lower state (in terms of potential energy).

If you buy a cup of hot chocolate and forget to drink it, you will discover that it has lost energy and the temperature has dropped to the temperature of the surrounding room.

You can explain the tendency of atoms to behave in much the same way by comparing the strength of the bonds that hold the atoms together in compounds (and also in polyatomic elements).

Example

Adding energy to methane gas molecules in the presence of oxygen molecules causes the energy to break the atoms apart from each other.

The individual atoms are now free to collide and form new bonds with atoms that are near to them. This could not happen until enough energy was available to break up the original molecules.
$CH_4 + O_2 + \text{energy} \rightarrow C + H + O$ (unstable)

Then, $C + O \rightarrow CO_2$ and $H + O \rightarrow H_2O$ (both more stable, with lower energy than CH_4)

If a collision results in new bonds with stronger attractions between the atoms than in the original compound (in this case, CH_4), then the new compounds (in this case, CO_2 and H_2O) are considered more stable, since it will be harder to break those bonds apart (decompose them) again. The system has reached a lower overall energy level and is chemically at its minimum energy.

ENTROPY

Entropy means disorder or chaos. *Higher entropy* means less organized. Often, non-chemical examples can help explain this; for example, if you build a brick wall and put a great deal of energy into your project, you have added "order" to the bricks. If someone then decides to break that wall, breaking it all apart, the bricks do not spontaneously rearrange themselves back into that wall.

Almost all chemical arrangements appear to follow that same set of rules. Molecules break apart and do not automatically rebuild themselves. Reactions are reversible to some extent when conditions change to "force" the reverse reaction, such as when two NO_2 molecules can form the more complex N_2O_4 molecule when they are in a pressurized system.

As a general rule, systems change over time to reach the minimum energy content (their most stable conditions) and the greatest number of separate particles. Sometimes these tendencies work against each other, as in the following example.

Example

Sulfur dioxide reacts with oxygen to form sulfur trioxide.
$2SO_2 + O_2 \rightleftharpoons SO_3 + \text{energy}$

The forward reaction releases energy, as the sulfur trioxide is more stable. However, the reverse reaction results in more molecules—an increase in entropy. If this happens in a closed system, a point is reached at which the two tendencies balance each other—a dynamic equilibrium is reached.

Practice

Use the following information to answer the next question.

The enthalpy and entropy changes that are accompanied in four-reaction systems are listed in the given table.

Reaction system	Enthalpy	Entropy
System 1	Increases	Increases
System 2	Decreases	Increases
System 3	Increases	Decreases
System 4	Decreases	Decreases

5. The formation of products is **most** favourable in which reaction system?
   A. System 1 **B.** System 2
   C. System 3 **D.** System 4

12.3.1.6 *describe, using the concept of equilibrium, the behaviour of ionic solutes in solutions that are unsaturated, saturated, and supersaturated*

SOLUTIONS AND EQUILIBRIUM

The solubility of a substance is the mass of solute to form a saturated solution in a given volume at a given temperature. Solutions that are unsaturated can still dissolve more solute; they are not at their maximum concentration. Solutions that are supersaturated have more solute than a saturated solution at the same temperature. This is an unstable solution in which precipitation of the solute is highly likely. A saturated solution is a solution that has its maximum concentration at a given temperature. Just because a solution is saturated does not mean no more solute will dissolve. The reason is that dissolving does not stop. Crystallization and dissolution are both occurring simultaneously. When the solution is saturated, the rates of dissolution and crystallization are equal. This is one type of *dynamic equilibrium*.

Practice

6. A saturated solution at equilibrium with an undissolved solute could be **best** described as a system in which
   A. dissolution and crystallization are equal
   B. the change in the reaction favours dissolution
   C. the change in the reaction favours crystallization
   D. competing processes are occurring that cannot balance each other

12.3.1.7 *define constant expressions, such as K_{sp}, K_w, K_a, and K_b*

EXPRESSING EQUILIBRIUM CONSTANT

The equilibrium constant expression is given in the general form

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

for the reaction $aA + bB \rightleftharpoons cC + dD$. However, there are many different types of equilibrium and consequently many different names for the equilibrium constant, although the general form is the same.

SOLUBILITY PRODUCT CONSTANT

SYMBOL

K_{sp}

DEFINITION

The maximum amount of ions that can be dissolved at a given temperature. The solid reactant is never included in the constant expression.

Example

$CaCO_{3(s)} \rightleftharpoons Ca^{2+}_{(aq)} + CO^{2-}_{3(aq)}$

$K_{sp} = [Ca^{2+}][CO_3^{2-}]$

ACID DISSOCIATION CONSTANT

SYMBOL

K_a

DEFINITION

Acids ionize into a hydrogen ion, H^+, and a conjugate base, A^-. For any acid, K_a is the dissociation of that acid at a given temperature.

$HA \rightleftharpoons H^+ + A^-$

$$K_a = \frac{[A^-][H^+]}{[HA]}$$

Example

$H_2CO_{3(aq)} \rightleftharpoons HCO^-_{3(aq)} + H^+_{(aq)}$

$$K_a = \frac{[HCO_3^-][H^+]}{[H_2CO_3]}$$

BASE DISSOCIATION CONSTANT

SYMBOL

K_b

DEFINITION

Bases gain protons (or hydrogen ions), creating hydroxide ions, OH^-, in solution and a conjugate acid, HB^+. For any base, K_b is the association constant for the protonation of that base at a given temperature.

$B + H_2O \rightleftharpoons HB^+ + OH^-$

$$K_b = \frac{[HB^+][OH^-]}{[B]}$$

Example

$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH^+_{4(aq)} + OH^-_{(aq)}$

$$K_b = \frac{[OH^-][NH_4^+]}{[NH_3]}$$

IONIZATION OF WATER

SYMBOL

K_w

DEFINITION

Water can act as an acid or a base, ionizing into hydrogen and hydroxide ions. The concentration of water is always constant; therefore, it is not included in the equilibrium expression. The self-ionization of pure water is a constant value.

$K_w = [OH^-][H_3O^+] = 1.0 \times 10^{-14}$

Example

$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$

$K_w = [OH^-][H_3O^+] = 1.0 \times 10^{-14}$

Practice

7. Which of the following expressions correctly represents the solubility product constant (K_{sp}) for the following dissolution reaction?

$Ag_2C_2O_{4(s)} \rightleftharpoons 2Ag^+_{(aq)} + C_2O^{2-}_{4(aq)}$

A. $K_{sp} = \dfrac{Ag_2C_2O_4}{[Ag^+]^2[C_2O_4^{2-}]}$

B. $K_{sp} = [Ag^+][C_2O_4^{2-}]^2$

C. $K_{sp} = \dfrac{[Ag^+]^2[C_2O_4^{2-}]}{[Ag_2C_2O_4]}$

D. $K_{sp} = [Ag^+]^2[C_2O_4^{2-}]$

12.3.1.8 *compare strong and weak acids and bases using the concept of equilibrium*

ACID/BASE STRENGTH AND K_C

Strong acids ionize completely (quantitatively) in water to form $H_3O^+_{(aq)}$ and a balancing anion (conjugate base). Weak acids ionize only partially (sometimes only very slightly) to form $H_3O^+_{(aq)}$ and a balancing anion. Strong bases either dissociate completely (quantitatively) in water to form $OH^-_{(aq)}$ and a conjugate acid or the completely protonated substrate, creating $OH_{(aq)}$ and a conjugate acid (cation). Weak bases react only partially (sometimes only very slightly) to form $OH^-_{(aq)}$ and a balancing species. Because of the greater ion concentration, strong acids have a lower pH, react more quickly, and have a greater conductivity than weak acids of the same concentration. A similar statement can be made for bases.

The relative strength of acids and bases can be observed in the quantitative K_a and K_b values. Given that $K_a = \dfrac{[A^-][H^+]}{[HA]}$ and

$K_b = \dfrac{[HB^+][OH^-]}{[B]}$,

it can be seen mathematically that strong acids and bases will have high concentrations of products and therefore high K_a and K_b values. Conversely, a weak acid or base will have low product concentrations and high reactant concentrations, thus giving rise to low $K_a$and K_b values. If an acid has a high K_a value (a strong acid), it will have a low K_b value (it will be a weak base). This is because $K_w = K_a \times K_b$, so $K_a \propto \dfrac{1}{K_b}$.

Practice

Use the following information to answer the next question.

The dissociation constants of hypoiodous acid (HIO), benzoic acid (C_6H_5COOH), hypochlorous acid (HClO), and hydrocyanic acid (HCN) are 2.3×10^{-11}, 6.5×10^{-5}, 3.2×10^{-8}, and 4.9×10^{-10}, respectively.

8. If the given acids are arranged in descending order from strongest to weakest, then the correct order will be

A. HCN > HClO > HIO > C_6H_5COOH

B. C_6H_5COOH > HClO > HIO > HCN

C. HCN > HIO > HClO > C_6H_5COOH

D. C_6H_5COOH > HClO > HCN > HIO

Use the following information to answer the next multipart question.

9. An acid–base reaction occurs between HSO_3^- and IO_3^-.

Part A

Open Response

Write the equation for the equilibrium that results.

Part B

Open Response

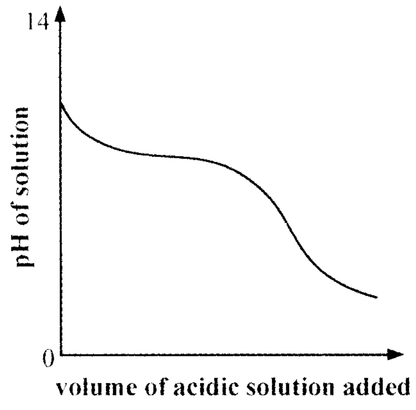
Identify one conjugate acid–base pair in the reaction.

12.3.1.9 *describe the characteristics and components of a buffer solution*

BUFFERS

A buffer is a solution that is able to resist changes in pH upon addition of small amounts of acidic or alkaline substances. Buffer solutions contain either a weak base and its conjugate acid or, more commonly, a weak acid and its conjugate base. On a titration curve, the buffer region is the "flattish" region of the graph.

Titration Curve



HOW DO BUFFERS RESIST pH CHANGE?

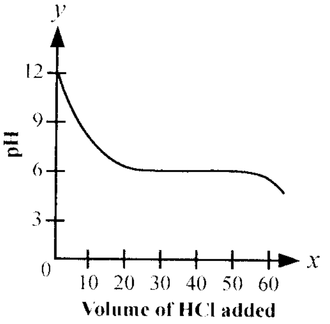
Buffer solutions can combine with both H^+ ions and OH^- ions depending on which "stress" is added. Each time more H^+ or OH^- ions are added, the equilibrium shifts to consume the added species (Le Châtelier's Principle) and a new equilibrium is established between the acid and its conjugate base. For example, in the case of weak acid HA, $HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A_{(aq)}$.

If a strong base is added, the amount of H_3O^+ (or H^+) will decrease, which would typically result in an increase in pH. However, the acid buffer equilibrium shifts left to produce more H_3O^+, thereby maintaining a similar pH level in the solution. If a strong acid were added, the conjugate base, A^-, would gain a proton, shifting the equilibrium to the right and again keeping the pH of the solution at a steady level.

Practice

Use the following information to answer the next question.

When a weak base is titrated with a strong acid, the pH of the solution initially decreases quickly and then decreases slowly until the end point is reached, as shown in the given graph.



10. Based on the given information, which of the following statements about the titration curve is correct?

A. The buffer region is alkaline.

B. The titration begins with a sharp increase in hydronium ion concentration.

C. The buffer region has a slope of zero.

D. The endpoint of the titration is at a pH of 6.

12.3.2.1 *use appropriate vocabulary to communicate ideas, procedures, and results related to chemical systems and equilibrium*

EQUILIBRIUM VOCABULARY

When communicating ideas and information, it is important to use the proper scientific vocabulary. The following chart lists terms that are frequently used in discussing chemical equilibrium. These terms are not discussed elsewhere in this unit.

Term	Definition
Closed system	A system in which heat may be exchanged with the surroundings but not matter (i.e., gas cannot escape).
Common ion or common ion effect	A solution that contains the same ion as the solute being dissolved. The "common ion" acts to decrease the solute solubility.
Conjugate acid–base pair	Two chemicals in an acid-base reaction that differ in the presence or absence of a proton (H^+). E.g., $HNO_{3(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + NO^-_{3(aq)}$ HNO_3 and NO_3^- are a conjugate acid–base pair and H_2O and H_3O^+ are a conjugate acid–base pair.
Electrolyte	A substance that contains free ions that can conduct electricity. E.g., $NaCl_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$
Heterogeneous system	A system in which species may exist in different states. E.g.,$Na_{(s)} + HCl_{(aq)} \rightleftharpoons NaCl_{(aq)} + H_{2(g)}$
Homogeneous system	A system in which species are all the same state. E.g., $H_{2(g)} + Cl_{2(g)} \rightleftharpoons 2HCl_{(g)}$

Practice

Use the following information to answer the next question.

$$CH_3COOC_2H_{5(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COOH_{(aq)} + C_2H_5OH_{(aq)}$$

11. Which of the following types of equilibrium occurs in the given reaction?

A. Static

B. Aqueous

C. Homogeneous

D. Heterogeneous

12.3.2.2 *apply Le Châtelier's principle to predict how various factors affect a chemical system at equilibrium, and confirm their predictions through experimentation*

LE CHÂTELIER'S PRINCIPLE

In 1884, the Frenchman Henri Louis Le Châtelier realized that, by changing the conditions (stresses) on a chemical system, the reactions involved could be forced to form a higher percentage of products or reversed to form a higher fraction of reactants. For chemical engineers, this was a major breakthrough, allowing them to control both the amount and rates of product formation. By varying concentrations, temperatures, and pressures, it is possible to maximize production but also to keep pressures and temperatures within safe operating limits.

A commonly cited example is the process developed by Fritz Haber in 1913.

Dr. Haber synthesized ammonia from hydrogen and nitrogen, originally using an iron oxide catalyst. He was able to significantly increase ammonia production by

- increasing pressure, encouraging the forward reaction
- removing ammonia as it was formed, also encouraging the forward reaction, which made up for the need to lower the temperature (which would have slowed down the forward reaction, decreasing ammonia production)

Simply stated, Le Châtelier's principle suggests that any change in conditions (the "stress") in a chemical equilibrium will result in an opposite shift.

TYPES OF STRESSES ON REACTIONS

HEAT

If heat is added to a system, the reaction will shift to use some of that added energy, partially reversing the process.

Example

What happens when energy is added to a sealed tube containing $NO_{2(g)}$ and $N_2O_{4(g)}$?

$NO_{2(g)} \rightleftharpoons N_2O_{4(g)} + \text{energy}$

There will be an increase in $[NO_{2(g)}]$ and a decrease in $[N_2O_{4(g)}]$ because the reverse reaction uses up some of the added energy.

If heat is removed, the opposing reaction will be increased, as the forward reaction produces energy.

CONCENTRATION

You can consider the effects of concentration changes by referring to the same equation. If you could add more $NO_{2(g)}$, increasing the concentration of $NO_{2(g)}$, the forward reaction would speed up, producing more $N_2O_{4(g)}$ and energy until a new balance of forward and reverse reactions is reached.

Decreasing the concentration of $NO_{2(g)}$ would shift the reaction to the left, as the $N_2O_{4(g)}$ uses up some energy to break apart into $NO_{2(g)}$, replacing some of the $NO_{2(g)}$ that was removed.

PRESSURE

Increasing pressure can be counteracted by the formation of larger molecules, in the same way that a shortage of downtown parking spaces can be eased by an increase in carpooling and public transit use. Think of the traffic congestion caused if a carpool of four people using one vehicle each day is replaced by each person driving alone in a vehicle.

In the sample equation, if pressure is increased (by decreasing volume), the stress can be eased by the formation of more $N_2O_{4(g)}$ because each one takes up approximately the same space as $2NO_{2(g)}$ molecules.

Decreasing pressure by increasing volume would cause the opposite effect: a left shift. There would be more room for the $NO_{2(g)}$ molecules, and there would be a smaller chance of collisions to form $N_2O_{4(g)}$ molecules.

Adding an inert gas does not change the concentrations of the substances involved in the equilibrium, so it does not shift the position of equilibrium.

EXPERIMENTAL VERIFICATION

A common demonstration of the effects of placing stresses on an equilibrium system involves a set of test tubes containing a solution of cobalt(II) ions dissolved in ethanol.

Six test tubes are placed in a test-tube rack, each one containing the same mixture of $CoCl_4^{2-}$, H_2O, $Co(H_2O)_6^{2+}$, and Cl^- dissolved in ethanol (eth) at room temperature. The mixture is purple because of the combination of blue $CoCl_4^{2-}$ and red $Co(H_2O)_6^{2+}$.

$$CoCl^{2-}_{4(eth)} + 6H_2O_{(eth)} \rightleftharpoons Co(H_2O)^{2+}_{6(eth)} + 4Cl^-_{(eth)} + \text{energy}$$

Example

Predict the colour changes in the solutions in terms of left and right shifts when the following stresses are placed on each of the tubes individually.

1. Add energy by placing in a warm-water bath
2. Remove energy by placing in an ice-water bath
3. Increase $[Cl^-]$ by adding a few NaCl crystals
4. Decrease $[Cl^-]$ by adding a few $AgNO_3$ crystals
5. Increase $[H_2O]$, using an eye dropper
6. Leave alone for comparison (colour changes)

The experimental results should resemble the following.

1. Becomes more blue, left shift (some of the red ion is removed by the added energy)
2. Becomes more red, right shift (reverse reaction is slowed by the removal of energy, forward reaction is still progressing until a new equilibrium is reached)
3. Cl^- ions react to use up red, therefore the solution becomes more blue as the reactant concentration increases
4. Ag^+ ions combine with Cl^- to form a precipitate, removing Cl^- from the solution, resulting in a shift right—the solution becomes more red
5. Blue is used up in the forward reaction, so the colour becomes more red as the concentrations of products increase
6. No colour change

Practice

Use the following information to answer the next question.

Methane can be obtained by the reaction of formaldehyde with water, as shown below.
$HCHO_{(g)} + H_2O_{(g)} \rightleftharpoons CH_{4(g)} + O_{2(g)}$

12. If the given reaction is assumed to be at equilibrium, then on the addition of CH_4, the equilibrium will

A. undergo no change

B. show minor random changes

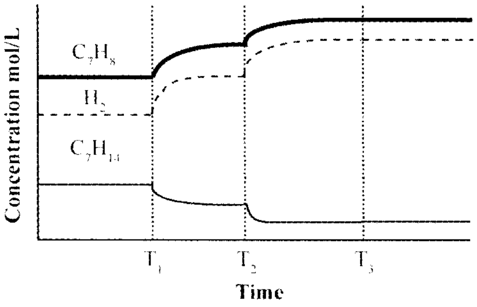
C. shift toward the forward direction

D. shift toward the backward direction

Open Response

Use the following information to answer the next question.

On the given graph is a representation of the equilibrium changes occurring to the following reaction:
energy + $C_7H_{14(g)} \leftrightharpoons C_7H_{8(g)} + 3H_{2(g)}$.



13. Suggest different stresses that may have caused the trends observed at times T_1, T_2, and T_3.

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12.3.2.3 *carry out experiments to determine equilibrium constants*

SOLVING EQUILIBRIUM PROBLEMS

Using the equilibrium constant expression

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b} \text{ for any reaction}$$

$aA + bB \leftrightharpoons cC + dD$, the equilibrium constant K_c can be determined for a given temperature. It is important to use only equilibrium concentrations of reactants and products in the expression.

Example

The following reaction takes place in catalytic methanation.
$3H_{2(g)} + CO_{(g)} \leftrightharpoons H_2O_{(g)} + CH_{4(g)}$

During an experiment, a student monitors this reaction in a closed 10.0 L vessel at 927°C until the pressure of the system no longer changes. At this point, the student determines the reaction to be at equilibrium and takes the necessary measurements to find the moles of each reactant and product. The findings are shown below.
What is the value of K_c for this reaction?

Step 1: Find the equilibrium concentration for each species.

$$[H_2]_{eq} = \frac{1.839 \text{ mol}}{10.0 \text{ L}} = 1.839 \text{ M}$$

$$[CO]_{eq} = \frac{0.631 \text{ mol}}{10.0 \text{ L}} = 0.0631 \text{ M}$$

$$[CH_4]_{eq} = \frac{0.387 \text{ mol}}{10.0 \text{ L}} = 0.0387 \text{ M} = [H_2O]_{eq}$$

Step 2: Plug equilibrium values into the K_c expression.

$$K_c = \frac{[CH_4][H_2O]}{[CO][H_2]^3} = \frac{(0.0387)(0.0387)}{(0.0613)(0.1839)^3} = 3.93$$

If the value of K_c is already given, equilibrium concentrations of reactants and products can be found. An ICE table is commonly used in these types of calculations to help keep track of changing concentrations.

Example

What is the concentration of each substance at equilibrium for the gaseous reaction $C_2H_{4(g)} + H_{2(g)} \leftrightharpoons C_2H_{6(g)}$, where $K_c = 0.99$ and the initial concentration of ethene and hydrogen are 0.305 M and 0.450 M, respectively?

Step 1: Write the K_c expression. $K_c = \dfrac{C_2H_6}{[H_2][C_2H_4]}$ $= 0.99$

Step 2: Make an ICE table.

	C_2H_4	+	H_2	$\leftrightharpoons$	C_2H_6
Initial	0.305		0.405		0
Change	$-x$		$-x$		$+x$
Eq	$0.305 - x$		$0.405 - x$		x

Step 3: Plug the equilibrium concentrations into the K_c expression.

$$K_c = 0.99 = \frac{x}{(0.305 - x)(0.405 - x)}$$

Step 4: Solve for x.

$$0.99 = \frac{x}{(0.305 - x)(0.405 - x)}$$

$$= \frac{x}{0.123 - 0.71x + x^2}$$

$$0 = 0.123 - 1.70x - x^2 \text{ or } x^2 + 1.70x - 0.123 = 0$$

Using the quadratic equation, the value for x can be found.

$$\frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-1.70 + -\sqrt{1.70^2 - 4x - 0.123}}{2}$$

$= 0.0695$ M or -1.77 M

Since -1.77 M is an impossible value (cannot have a negative concentration for C_2H_6), then the correct value is $x = 0.0695$ M.

Step 5: Find all equilibrium concentration values.

$[C_2H_4]_{eq} = 0.305 - 0.0695 = 0.236$ M

$[H_2]_{eq} = 0.405 - 0.0695 = 0.336$ M

$[C_2H_6]_{eq} = 0.0695$ M

These five basic steps are common for all types of equilibrium problems in which K_c is given.

The questions differ in how to solve for x in step 4, depending on the type of equation that results. Some will require the quadratic equation but others will be much simpler or will involve assumptions in which K_c is a very small value. For example, if in the above question, K_c were equal to 9.9×10^{-7}, then an assumption could be made in step 4.

Assume $0.305 - x \cong 0.305$ and $0.405 - x \cong 0.405$.

Check the assumption with the 100 Rule.

If $\dfrac{[\text{lowest concentration}]}{K_c} > 100$, the assumption is correct.

$\dfrac{0.305}{9.9 \times 10^{-7}} > 100$, thus the assumption is correct.

This assumption would then eliminate the need for the quadratic formula and simplify the equation. Such an assumption is only valid when the 100 Rule applies (K_c must be very small).

$n_{H_2} = 1.839$ mol; $n_{CO} = 0.631$ mol;

$n_{CH_4} = 0.387$ mol; $n_{H_2O} = 0.387$ mol

Practice

Use the following information to answer the next question.

$H_2 + I_2 \rightleftharpoons 2PHI$

14. In the given chemical reaction, what will be the equilibrium constant of the backward reaction if the equilibrium constant of the forward reaction is K?

A. $-K$ **B.** $\dfrac{1}{K}$

C. K **D.** K^2

Open Response

Use the following information to answer the next question.

Calcium nitrate and sodium hydroxide react to form an insoluble compound, calcium hydroxide. In an experiment, 750 mL of each reactant was combined in successively stronger concentrations in order to determine the K_{sp} of calcium hydroxide. The table below gives one student's results.

Concentration of Reactants in M

Concentration	Observation
0.002	clear
0.04	slightly foggy
0.1	some white precipitate
0.2	lots of white precipitate

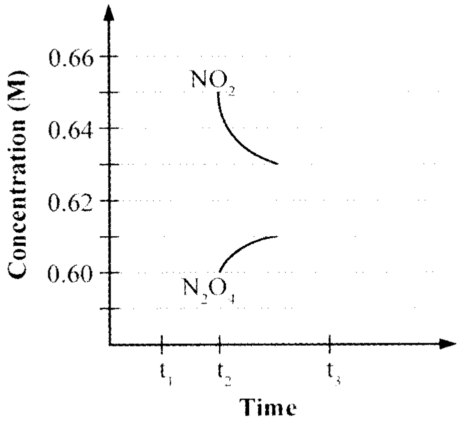
15. Find the solubility product constant for calcium hydroxide using the proper equilibrium constant expression for the dissolution of calcium hydroxide. You may need to complete part or all of a data table to find K_{sp}.

Open Response

16. Consider the question in which you found the K_{sp} of calcium hydroxide from experimental data. Would you expect the calculated value from the experiment to be higher or lower than the theoretical K_{sp} of $Ca(OH)_{2(s)}$? Support your answer and suggest a method that would give a more accurate value.

Use the following information to answer the next multipart question.

17. Consider the following diagram for the equilibrium $N_2O_{4(g)} \rightleftharpoons 2NO_{2(g)}$ $\Delta H = +59$ kJ.



Part A

Open Response

Calculate the value of K_{eq} at t_1.

Part B

Open Response

Is the K_{eq} value you found in the previous section of this question the true K_{eq} for this reaction? Why or why not?

12.3.2.4 *calculate the molar solubility of a pure substance in water or in a solution of a common ion, given the solubility product constant (K_{sp}), and vice versa*

DISSOCIATION EQUILIBRIUM

Ionic compounds usually dissociate into ions when they dissolve in water. If a slightly soluble compound, like AgCl, exists in excess in a saturated solution, an equilibrium develops between the solid compound and the dissolved ions. So for AgCl, the equilibrium is as follows:

$AgCl_{(s)} \leftrightharpoons Ag^{+}_{(aq)} + Cl^{-}_{(aq)}$

The K_{sp} expression is $K_{sp} = [Ag^{+}][Cl^{-}]$. Recall that solids do not change in concentration and so are not included in the K_{sp} expression. The molar solubility (maximum concentration of ions) for any slightly soluble compound can be determined knowing K_{sp} and the K_{sp} expression.

Example

If the K_{sp} for silver chloride is 1.77×10^{-10} at 25°C, what is the molar solubility of AgCl?

$K_{sp} = [Ag^{+}][Cl^{-}] = 1.77 \times 10^{-10}$

$[AgCl] = [Ag^{+}] = [Cl^{-}] = x$

$1.77 \times 10^{-10} = x^{2}$, so

$x = \sqrt{1.77 \times 10^{-10}} = 1.33 \times 10^{-5}$.

Therefore, the molar solubility of AgCl is 1.33×10^{-5} M.

COMMON ION EFFECT

The solubility of a slightly soluble compound can be decreased by the presence of a common ion in correlation with Le Châtelier's principle.

Example

$AgCl_{(s)} \leftrightharpoons Ag^{+}_{(aq)} + Cl^{-}_{(aq)}$

If a solution of AgCl is added to a solution of 0.1 M NaCl, the equilibrium will shift to the right to use up excess Cl^{-} ions, thereby decreasing the solubility of AgCl.

An ICE table is sometimes used in these types of calculations.

	AgCl	⇋	**Ag⁺**	+	**Cl⁻**
Initial			0		0.1
Change			$+x$		$+x$
Eq			x		$0.1 + x$

$K_{sp} = [Ag^{+}][Cl^{-}] = 1.77 \times 10^{-10}$

$[AgCl] = [Ag^{+}] = [Cl^{-}]$

$1.77 \times 10^{-10} = x \times (x + 0.1)$

Because silver chloride is only slightly soluble (K_{sp} is small), x is assumed to be negligible.

Assume $0.1\ M + x \cong 0.1$ M.

Check with the 100 Rule.

$\frac{0.1}{1.77 \times 10^{-10}} >> 100$, so the assumption is correct.

$1.77 \times 10^{-10} = x \times (0.1) = 0.1x$

$x = \frac{1.77 \times 10^{-10}}{0.1} = 2.0 \times 10^{-9}$

So the molar solubility of AgCl in the presence of 0.1 M NaCl is 2.0×10^{-9} M.

Practice

Use the following information to answer the next question.

The given equilibrium is obtained when $BaSO_4$ is added to water.

$BaSO_{4(s)} \rightleftharpoons Ba^{2+} + SO_4^{2-}$

18. Which of the following factors will **not** affect the solubility of $BaSO_4$ in water?

A. Increase in temperature

B. Decrease in pressure

C. Addition of $CaSO_4$

D. Addition of H_2SO_4

Numerical Response

Use the following information to answer the next question.

$AgCl_{(s)} \rightleftharpoons Ag^{+}_{(aq)} + Cl^{-}_{(aq)}$

19. If the K_{sp} for the given reaction is 1.8×10^{-10} at 25°C, what is the equilibrium concentration of chloride ions, $[Cl^{-}]_{eq}$, expressed in 10^{-5} M?

Open Response

20. The molar solubility of calcium iodiate, $Ca(IO_3)_2$, is 1.79×10^{-3} mol / L. What is the K_{sp} of this compound?

Open Response

Use the following information to answer the next question.

$CuI\ K_{sp} = 1.27 \times 10^{-12}$
$CuBr\ K_{sp} = 6.27 \times 10^{-9}$
$CuCl\ K_{sp} = 1.72 \times 10^{-7}$
$CuCN\ K_{sp} = 3.47 \times 10^{-20}$

21. Arrange the given compounds in order of increasing $[Cu^{+}]_{(aq)}$. Give the reasoning for your answer.

12.3.2.5 *predict the formation of precipitates by using the solubility product constant*

PREDICTING A PRECIPITATE

The solubility product constant can be used to determine whether a precipitate of the compound in question will occur. A trial ion product (Q) is calculated the same way K_{sp} is, and the value is compared to the actual K_{sp} of that compound to see if a precipitate will form.

- If $Q > K_{sp}$, the solution is supersaturated and a precipitate will form.
- If $Q = K_{sp}$, the solution is saturated and no precipitate will form.
- If $Q < K_{sp}$, the solution is unsaturated and no precipitate will form.

Example

Will a precipitate form if 100.0 mL of 0.0010 M $Pb(NO_3)_2$ are added to 100.0 mL of 0.0020 M $MgSO_4$? The K_{sp} of $PbSO_4 = 6.3 \times 10^{-7}$.

Step 1: Write the net ionic equation for the precipitate.

$PbSO_{4(s)} \leftrightharpoons Pb^{2+}_{(aq)} + SO^{2-}_{4(aq)}$

Step 2: Determine the concentrations of the ions involved in the **total combined volume**.

$$[Pb^{2+}] = [Pb(NO_3)_2]$$

$$= \frac{C_1V_1}{V_2} = \frac{0.0010\ M \times 0.1\ L}{0.2\ L} = 5.0 \times 10^{-4}\ M$$

$$[SO_4^{2-}] = [MgSO_4] = \frac{C_1V_1}{V_2} = \frac{0.0020\ M \times 0.1\ L}{0.2\ L}$$

$$= 1.0 \times 10^{-3}\ M$$

Step 3: Write the equilibrium constant expression and find Q.

$$Q = [Pb^{2+}][SO_4^{2-}]$$

$$= (5.0 \times 10^{-4}) \times (1.0 \times 10^{-3})\ M = 5.0 \times 10^{-7}$$

Step 4: Compare Q with the K_{sp} to determine if a precipitate results.

$Q < K_{sp}$, therefore no precipitate will form.

Practice

22. Two salt solutions were mixed and a trial K_{sp} or Q (reactant quotient) was calculated to be 2.0×10^{-9}. The K_{sp} value is 1.0×10^{-10}. From this information, which of the following charts is correct?

A.

K_{sp} Comparison	**Outcome**
$Q < K_{sp}$	Precipitate forms

B.

K_{sp} Comparison	**Outcome**
$Q > K_{sp}$	Precipitate forms

C.

K_{sp} Comparison	**Outcome**
$Q < K_{sp}$	No precipitate forms

D.

K_{sp} Comparison	**Outcome**
$Q > K_{sp}$	No precipitate forms

Use the following information to answer the next multipart question.

23. A sufficient amount of $Na_2SO_{4(aq)}$ is added to 0.1 M $BaCl_2$ to cause precipitation.

Part A

Open Response

Write the net ionic equation for the precipitate formed.

Part B

Open Response

Calculate the $[SO_4^{2-}]$ at the moment the precipitate starts forming if the $K_{sp} = 1.1 \times 10^{-10}$.

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12.3.2.6 *solve equilibrium problems involving concentrations of reactants and products and the following quantities: K_{eq}, K_{sp}, K_a, K_b, pH, pOH*

EQUILIBRIUM AND pH

Many problems involving equilibrium demonstrate the relationship between water, acids, and bases. The use of pH in some of the questions aids in representing the concentrations of the solutes within the solution, and the acidity or alkalinity of that solution. Where the acid and base dissociation constants (K_a and K_b) are used, pH can also be included either as the target unknown of the question or as a given measurement.

$$K_a = \frac{[A^-][H^+]}{[HA]} \text{ and } K_b = \frac{[HB^+][OH^-]}{[B]}$$

From the equilibrium constant expressions above, pH can be determined using $[H^+]$ and $[OH^-]$ values. For the K_a expression, the equilibrium concentration of H^+ or $[H^+]_{eq}$ is plugged into the pH equation.

$\text{pH} = -\log[H^+]$

For basic solutions, pH is used in one of two ways:

- Use $K_w = [OH^-][H^+] = 1.0 \times 10^{-14}$ to find $[H^+]$ and then pH.
- Find pOH
  $\text{pH} = -\log[OH^-]$ and then use $14 = \text{pH} + \text{pOH}$ to find pH.

Practice

24. If the concentration of the reactants is doubled in a reversible reaction, then the equilibrium constant

A. is halved

B. is doubled

C. is quartered

D. remains the same

Numerical Response

Use the following information to answer the next question.

Coal and natural gas contain trace amounts of sulfur compounds, which, when burned, may lead to acid rain pollution.

Reactions Related to Acid Rain

I. $2H_2S_{(g)} + 3O_{2(g)} \rightleftharpoons 2H_2O_{(g)} + 2SO_{2(g)}$

II. $2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$

III. $SO_{2(g)} + H_2O_{(l)} \rightleftharpoons H_2SO_{3(aq)}$

IV. $2SO_{3(g)} + 2H_2O_{(l)} \rightleftharpoons 2H_2SO_{4(aq)}$

At 900 K, the equilibrium constant for reaction II is 13.0. The equilibrium concentrations are

$[SO_{2(g)}] = 0.361 \text{ mol/L}$

$[SO_{3(g)}] = 0.840 \text{ mol/L}$

25. Taking into account the given values, calculate the equilibrium concentration of $O_{2(g)}$ ____.

Numerical Response

Use the following information to answer the next question.

Methanoic (formic) acid is the irritant secreted during an ant bite. The irritation is partially the result of the ionization of methanoic acid. The equilibrium equation for the ionization can be represented as $HCOOH_{(aq)} + H_2O_{(l)}$

$\rightleftharpoons H_3O_{(aq)} + HCOO^-_{(aq)}$

26. The pH of 0.10 mol/L$NaHCOO_{(aq)}$ is ____. (Record your **three-digit** answer.)

Open Response

Use the following information to answer the next question.

Consider the following equilibrium:
$CO_{(g)} + H_2O_{(g)} \rightleftharpoons CO_{2(g)} + H_{2(g)}$
$K_{eq} = 4.6$

27. Initially, 0.50 mol CO, 0.50 mol H_2O, 0.62 mol CO_2, and 0.62 mol H_2 are placed in a 1.0 L container, and the reaction proceeds toward products. Calculate the equilibrium $[H_2]$.

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12.3.2.7 *predict, in qualitative terms, whether a solution of a specific salt will be acidic, basic, or neutral*

SALT SOLUTIONS

Salts are ionic compounds consisting of an anion (– ion) and a cation (+ ion). Since salts are the products of acid/base reactions, salts can also be considered to be the combination of the anion of an acid and the cation of a base. The pH of a salt dissolved in water depends on the anions and cations of which it is composed.

Example

$NaHCO_{3(s)} + H_2O_{(l)} \rightleftharpoons Na^+_{(aq)} + H_2CO_{3(aq)} + OH^-_{(aq)}$

Sodium bicarbonate is the combination of NaOH (a strong base) and H_2CO_3 (a weak acid). As seen in the reaction above, the weak acid anion $HCO^-_{3(aq)}$ readily combines with free protons in the water to form $H_2CO_{3(aq)}$. This causes a decrease in H^+ ions and thus an increase in pH, making the solution basic.

GENERAL RULES OF SALTS

- Strong acid anion and weak basic cation = acidic solution
- Strong acid anion and strong basic cation = neutral solution
- Weak acid anion and strong basic cation = basic solution

It is often difficult to determine the pH of solutions with weak acid anions and weak base cation salts, as it often depends on which is weaker.

Practice

28. Which of the following salts is acidic in nature?

A. Na_2SO_4

B. Na_2S

C. Na_2SO_3

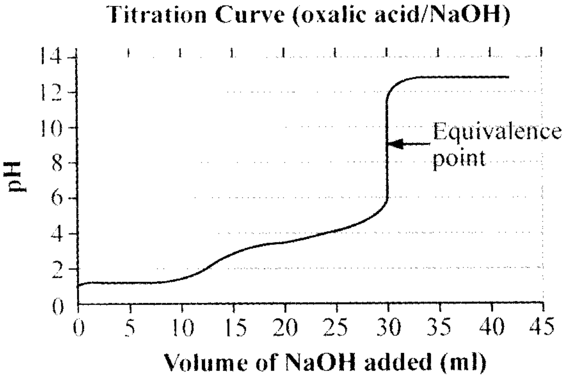
D. $NaHSO_3$

12.3.2.8 *solve problems involving acid-base titration data and the pH at the equivalence point*

EQUIVALENCE POINTS

The equivalence point in an acid/base titration is the point in the reaction at which equal stoichiometric amounts of acids and bases have been reacted. All the initial reactants have reacted and what is left is the resulting salt and water. On a graph, this is the vertical section of the graph, especially in a strong acid/strong base titration.

Equivalence Point on Titration Curve



Example

If 50 mL of 0.20 M ammonia, $NH_{3(aq)}$, is titrated by 0.12 M hydrochloric acid, what is the pH at the equivalence point for the resulting solution? K_b for ammonia is 1.8×10^{-5}.

Step 1: Determine the total volume of the solution.

$V_T = V_b + V_a = 50\text{ mL} + V_a$

$$V_a = \frac{0.1\text{ mol/L} \times 0.050\text{ L}}{0.12\text{ mol/L}} = 0.042\text{ L}$$

So, $V_T = 0.05\text{ L} + 0.042\text{ L} = 0.092\text{ L}$.

Step 2: Write the equilibrium equation at the equivalence point.

At the equivalence point, it can be assumed the acid–base reaction is complete and that only the salt and the water are left.

$NH_4Cl_{(aq)} \leftrightharpoons NH^+_{(aq)} + Cl^-_{(aq)}$

Cl^- ions are from a strong acid and will not combine with H^+ from the water. NH_4^+ ions are from a weak base and will dissociate in the water.

$NH^+_{4(aq)} + H_2O_{(l)} \leftrightharpoons NH_{3(aq)} + H_3O^+_{(aq)}$

This increases the concentration of H_3O^+ or H^+ in the water, so the expected pH value should be lower than 7.

Step 3: Find C_{salt} at equilibrium.

$$C_{\text{salt}} = [NH_4^+] = \frac{0.05\text{ L} \times 0.2\text{ mol/L}}{0.092\text{ L}} = 0.11\text{ M}$$

Step 4: Use K_b or K_a to find the $[H_3O^+]$.

$$K_w = K_a \times K_b\text{, so } K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}}$$

$$= 5.6 \times 10^{-10}$$

$$K_a = \frac{[NH_3][H^+]}{[NH_4^+]} = 5.6 \times 10^{-10}$$

	NH_4^+	$\leftrightharpoons$	NH_3	+	H^+
Initial	0.11		0		≅ 0
Change	$-x$		$+x$		$+x$
Eq	$0.11 - x$		x		x

$$K_a = 5.6 \times 10^{-10} = \frac{x^2}{0.11 - x}$$

Because K_a is so small, assume $0.11 - x \cong 0.11$.

Check with the 100 Rule: $\dfrac{0.11}{5.6 \times 10^{-10}} >> 100$, therefore the assumption is correct.

$$K_a = 5.6 \times 10^{-10} = \frac{x^2}{0.11}\text{ so } x^2 = 6.16 \times 10^{-11}$$

$$x = \sqrt{6.16 \times 10^{-11}} = 7.8 \times 10^{-6}$$

So $[H_3O^+] = x = 7.8 \times 10^{-6}\text{ M}$

Therefore, the pH of the solution at equilibrium is $\text{pH} = -\log(7.8 \times 10^{-6}) = 5.1$.

The following is the general format for solving questions dealing with the pH of a solution at an equivalence point.

When determining the pH of a solution at an equivalence point, you are finding the pH of the salt solution that is left. Whether the pH will be acidic or basic depends on the acid and base that initially formed the salt.

Practice

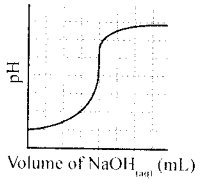
Use the following information to answer the next question.

Poisonous oxalic acid is found in non-toxic concentrations in vegetables such as spinach and rhubarb. Manufacturers of spinach juice are required to analyse the concentration of oxalic acid to avoid problems that could arise from unexpectedly high concentrations of oxalic acid. The reaction of oxalic acid with acidified potassium permanganate can be represented by the following equation.

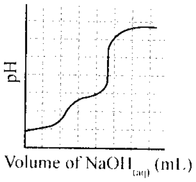
$$5HOOCCOOH_{(aq)} + 2MnO^-_{4(aq)} + 6H^+_{(aq)} \rightarrow 2Mn^{2+}_{(aq)} + 8H_2O_{(l)} + 10CO_{2(g)}$$

29. When oxalic acid is titrated with $NaOH_{(aq)}$, the titration curve that would be predicted is

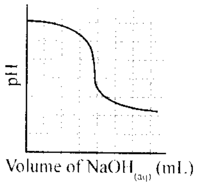
A.



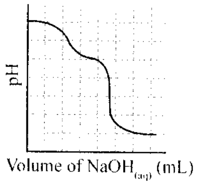
B.



C.



D.



Use the following information to answer the next multipart question.

30. A technician performed a titration to determine the concentration of a 27.0 mL sample of $NaOH_{(aq)}$. A few drops of phenol red indicator were added to the base, which was then titrated with a 0.24 mol/L solution of $HCl_{(aq)}$ until the indicator changed colour from red to orange.

Volume of Acid Used

Final buret reading (mL)	25.8
Initial buret reading (mL)	7.8

Part A

Numerical Response

The concentration of the $NaOH_{(aq)}$ solution was ____ mol/L. (Record your **three-digit** answer.)

Part B

Open Response

Using the data supplied, draw a curve for a sample of hydrochloric acid titrated with sodium hydroxide. Label the equivalence point and identify a suitable indicator.

12.3.3.1 *explain how equilibrium principles may be applied to optimize the production of industrial chemicals*

12.3.3.3 *explain how buffering action affects our daily lives, using examples*

12.3.3.2 *identify effects of solubility on biological systems*

APPLICATIONS OF CHEMICAL EQUILIBRIUM

The applications of chemical equilibrium are widespread, spanning many industrial, biological, and ecological uses. Industrial production of certain chemicals such as ammonia (commonly used in fertilizers) or sulfuric acid (used in ore processing, manufacturing of fertilizers, etc.) is big business, generating billions of dollars every year. Maximum production of industrial chemicals is made possible by applying the concepts of Le Châtelier's principle.

Solubility principles and the solubility constant are the foundation for understanding how many medical practices developed. The solubility of gases in blood for scuba divers and vapour anaesthetics in patients must be fully understood to ensure the safety of those involved. In addition, the precepts of solubility are applied to the use of barium sulfate in medical X-ray diagnosis and in the preventative treatment of kidney stones by increasing calcium solubility in urine.

Buffers and pH have several biological functions, including

- natural buffers in blood that keep blood pH within a narrow range
- commercial buffers in fruit juices to regulate taste
- the production of many fermented products such as cheese or yogurt

Practice

Use the following information to answer the next multipart question.

31. Sulfuric acid, $H_2SO_{4(aq)}$, has many uses, some of which include the production of fertilizers, removal of rust (oxides) from iron and steel, as an electrolyte in batteries, and even as a treatment for some types of heart disease. Since it is so widely employed, it is no surprise that the production of sulfuric acid is big business and manufacturers want to produce as much as possible. The process for producing sulfuric acid is called "the contact process" and can be summarized as follows:

$S_{8(s)} + 8O_{2(g)} \rightleftharpoons 8SO_{2(g)} \Delta H^\circ = -2\ 374 \text{ kJ}$

$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)} \Delta H^\circ = -198 \text{ kJ}$

(Catalyst added to this step)

$SO_{3(g)} + H_2O_{(l)} \rightleftharpoons H_2SO_{4(aq)}$

$\Delta H^\circ = -226 \text{ kJ}$

Part A

Open Response

Since an optimal yield of sulfuric acid means more money, industries are always trying to improve their production of the acid. They want as much sulfuric acid as possible as quickly as possible. In a concise paragraph, answer the following questions.

Evaluate the effect of temperature on the production of sulfuric acid. Use Le Châtelier's principle and your knowledge of chemical kinetics in your answer.

Part B

Open Response

Which reaction is the rate determining step? How might the rate and yield of this reaction be increased?

Part C

Open Response

Industrial plants that produce sulfuric acid also generate another product that is highly advantageous. What is this product and why is it so beneficial?

Use the following information to answer the next multipart question.

32. Many barium salts are highly poisonous. They are often used in pest control, such as in rat poisons, and can be harmful to humans if ingested. However, barium sulfate is used in medical diagnostics, where it creates more contrast of the internal organs so they are more clearly seen in an X-ray. Patients are often required to drink large volumes of barium sulfate prior to a gastrointestinal X-ray exam, and they suffer no ill side effects.

Part A

Open Response

Why is barium sulfate $\left(K_{sp} = 1.1 \times 10^{-10}\right)$ safe to ingest, while other barium salts such as barium nitrate $\left(K_{sp} = 4.64 \times 10^{-3}\right)$ are highly poisonous?

Part B

Open Response

Explain the use of magnesium sulfate as a treatment for barium poisoning using a chemical equation.

Use the following information to answer the next multipart question.

33. A buffer is any solution that has the ability to resist changes in pH when small amounts of acids or bases are added to it. Many important bodily functions are dependent on the buffering action of blood. Any significant change in pH could result in the death of an organism.

Part A

Open Response

Using the Internet or other resources, research the answers to the following questions and report your findings in a concise paragraph.

Give the bicarbonate acid/base equilibrium that exists in blood to act as a buffer.

Part B

Open Response

Name two vital functions of blood that are dependent on a controlled pH level.

Part C

Open Response

What normal bodily processes commonly change the pH in blood, and how does the body compensate for these changes?

SOLUTIONS–CHEMICAL SYSTEMS AND EQUILIBRIUM

1. **A**
2. **D**
3. **B**
4. **OR**
5. **B**
6. **A**
7. **D**
8. **D**
9. Part A- **OR**
   Part B- **OR**
10. **B**
11. **C**
12. **D**
13. **OR**
14. **B**
15. **OR**
16. **OR**
17. Part A- **OR**
    Part B- **OR**
18. **B**
19. **1.3**
20. **OR**
21. **OR**
22. **B**
23. Part A- **OR**
    Part B- **OR**
24. **D**
25. **0.416**
26. **8.38**
27. **OR**
28. **D**
29. **B**
30. Part A- **0.16**
    Part B- **OR**
31. Part A- **OR**
    Part B- **OR**
    Part C- **OR**
32. Part A- **OR**
    Part B- **OR**
33. Part A- **OR**
    Part B- **OR**
    Part C- **OR**

1. A

Dynamic equilibrium occurs when, in a closed system, a reversible reaction proceeds constantly in the forward and reverse directions such that the system's macroscopic properties are constant. A closed system cannot exchange matter with its surroundings. Thus, methane combustion in a Bunsen flame is not at equilibrium. (**Note:** No catalyst is present, the temperature is probably constant, and the state of the reactants and products is irrelevant.)

2. D

At equilibrium, the rate of the forward reaction will become equal to the rate of the backward reaction.

3. B

According to Le Châtelier's principle, with the decrease in the concentration of the products, the reaction proceeds in the forward direction. Therefore, in this case, the concentration of hydrogen gas will increase.

4. Open Response

1. Entropy increases because the change of state reflects a greater disorder of the molecules (there is a solid in the reactants but no solid in the products).
2. There are fewer large molecules in the reactants but many small molecules in the production, which also indicates an increase in disorder or entropy.

5. B

The formation of products is most favourable for the system that is accompanied by a decrease in enthalpy and an increase in entropy. Both of these factors are responsible for the spontaneity of a reaction.
In reaction system 2, enthalpy is decreasing and entropy is increasing. Hence, the formation of products is most favourable in this system.

6. A

During solution equilibria, the rates of forward and reverse processes (dissolution and crystallization, respectively) are equal.

7. D

The solubility product constant is found by multiplying the square of the silver ion concentration and the concentration of $C_2O_4^{2-}$ ions.

8. D

The greater the dissociation constant of an acid, the greater its acidic strength.
Therefore, arranged in descending order from strongest to weakest, the correct order of the strengths of the given acids will be
$C_6H_5COOH > HClO > HCN > HIO$.

9. Part A – Open Response

$HSO_3^- + IO_3^- \rightleftharpoons SO_3^{2-} + HIO_3$ (equilibrium arrow required)

Part B – Open Response

HSO_3^- and SO_3^{2-} OR IO_3^- and HIO_3

10. B

The best answer to this question is that the pH initially drops very rapidly as a result of the increase in the concentration of hydronium ions relative to the number of hydroxyl ions in the solution. Although the curve "flattens out," no titration curve actually reaches a slope of zero. B is the correct answer.

11. C

Since the reactants and products are in the same phase (homogeneous solution phase), the equilibrium is homogeneous.

12. D

According to Le Châtelier's principle, on increasing the amount of the products, the reaction will proceed in such a direction so as to undo the change imposed, i.e., shift toward the backward direction.

13. Open Response

T_1 and T_2 both show stresses that would cause a gradual shift toward the products. An increase in temperature and a decrease in pressure of the system would cause the equilibrium to shift to the right toward the products, giving the trends seen at T_1 and T_2.

At T_3, nothing is observed to happen. The only thing that may have been added is a catalyst, as catalysts do not affect the position of the equilibrium.

14. B

The equilibrium constant is proportionally constant. Every reaction has a fixed equilibrium constant. If the reaction is reversed, the value of the equilibrium constant will be reversed.

15. Open Response

Only the values of the "slightly foggy" trial are needed (when the concentration of reactants are 0.04 M), but the entire table is filled in for your benefit.

The table below shows the concentration of the reactants to the concentration of the products.

Reactants (M)	Products (M)
0.002	1.0×10^{-3}
0.04	0.02
0.1	0.05
0.2	0.1

$$[Ca^{2+}] = C_2 = \frac{C_1V_1}{V_2} = \frac{0.04\text{ M} \times 0.75\text{ L}}{1.5\text{ L}} = 0.02\text{ M}$$

$$Ca(OH)_{2(aq)} \leftrightharpoons Ca^{2+}_{(aq)} + 2OH^-_{(aq)}$$

$$K_{sp} = [Ca^{2+}][OH^-]^2$$

Use the trial in which the solution is seen as foggy because that is the first sign that a precipitate is forming.

$$K_{sp} = (0.02)(0.02)^2 = 8.0 \times 10^{-6}$$

16. Open Response

The experimental K_{sp} would be higher than the theoretical value (5.5×10^{-6}). This is because in the trial that was used to determine K_{sp}, a foggy precipitate was already forming, meaning the concentrations of NaOH and $Ca(NO_3)_2$ are too high. To rectify this error in the future, smaller concentration changes are needed and perhaps improved methods of observation, such as a magnifying glass or placing the beaker on a dark surface so that the foggy precipitate is seen sooner.

17. Part A – Open Response

$$K_{eq} = \frac{[NO_2]^2}{[N_2O_4]} = \frac{(0.65)^2}{0.60} = 0.70$$

Part B – Open Response

The K_{eq} that was found previously is not the true K_{eq} for the reaction. The equilibrium constant must be calculated by using concentrations found when the reaction has reached equilibrium. The equilibrium constant at t_1 is only true for that particular time, and it acts as a measure of the actual equilibrium as the reaction can still proceed further. Therefore, the K_{eq} must be calculated when the concentrations are constant (no net change with time) or after t_3. $K_{eq} = \frac{(0.63)^2}{0.61} = 0.65$

18. B

The effect of pressure is observed only in the case of gases. An increase in pressure increases the solubility of a gas in a liquid. In the present case, a solid substance is dissolved in a liquid. Hence, a change in pressure will have no effect on the solubility of $BaSO_4$ in water.

19. 1.3

$$K_{sp} = 1.8 \times 10^{-10} = [Ag^+][Cl^-] = x^2$$

$$x = \sqrt{1.8 \times 10^{-10}}$$

$$x = 1.3 \times 10^{-5}\text{ M}$$

Therefore, $[Cl^-] = 1.3 \times 10^{-5}$ M.

20. Open Response

$Ca(IO_3)_{2(s)}$	$\leftrightharpoons$	$Ca^{2+}_{(aq)}$	+	$2IO^-_{3(aq)}$
x		x		$2x$

$$K_{sp} = [Ca^{2+}][IO_3^-]^2 = (1.79 \times 10^{-3})((2)1.79 \times 10^{-3})^2$$
$$= 2.29 \times 10^{-8}$$

21. Open Response

The higher the K_{sp} value, the more soluble the compound, and therefore the greater the concentration of $Cu^+_{(aq)}$. Thus, in order of increasing solubility, the compounds should be CuCN, CuI, CuBr, and CuCl.

22. B

If the reactant quotient, Q, is greater than the reference K_{sp}, a precipitate will form. 2.0×10^{-9} is greater than 1.0×10^{-10}, therefore a precipitate will form.

23. Part A – Open Response

The formula equation for the reaction that occurs can be written as

$BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightleftharpoons BaSO_{4(s)} + 2NaCl_{(aq)}$.

The ionic equation is

$Ba^{2+}_{(aq)} + 2Cl^{-}_{(aq)} + 2Na^{+}_{(aq)} + SO^{2-}_{4(aq)}$

$\rightleftharpoons BaSO_{4(s)} + 2Na^{+}_{(aq)} + 2Cl^{-}_{(aq)}$.

Hence, the net ionic equation is represented as

$Ba^{2+}_{(aq)} + SO^{2-}_{4(aq)} \rightarrow BaSO_{4(s)}$.

Part B – Open Response

$$K_{sp} = [Ba^{2+}][SO_4^{2-}]$$

$$[SO_4^{2-}] = \frac{1.1 \times 10^{-10}}{0.1}$$

$$= \frac{1.1 \times 10^{-10}}{0.1}$$

$$[SO_4^{2-}] = 1.1 \times 10^{-9}\ M$$

24. D

The equilibrium constant is affected only by a change in the temperature. Hence, the equilibrium constant of the reaction remains the same.

25. 0.416

The equilibrium constant expression for reaction II is

$$K_c = \frac{[SO_{3(g)}]^2}{[SO_{2(g)}]^2[O_{2(g)}]}.$$

When rearranged for $[O_{2(g)}]$,

$$[O_{2(g)}] = \frac{[SO_{3(g)}]^2}{[SO_{2(g)}]^2\ K_{eq}}$$

Substitute the given values to get

$$[O_{2(g)}] = \frac{(0.840\ mol/L)^2}{(0.361\ mol/L)^2(13.0\ mol/L)}\cdot$$

$$= 0.416\ mol/L$$

26. 8.38

The K_b of $HCOO^-_{(aq)}$ can be calculated from the K_a of formic or methanoic acid. According to the table of relative strengths of acids and bases, the K_a is

1.8×10^{-4}, so the K_b of $HCOO^-_{(aq)}$ is therefore

$$K_b = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-4}} = 5.6 \times 10^{-11}.$$

The ICE table would appear as follows:

$HCOO^-_{(aq)}$ +	$H_2O_{(l)} \rightleftharpoons$	$OH^-_{(aq)}$ +	$HCOOH_{(aq)}$
$0.1 - x$		x	x

The concentration of hydroxyl ions can be found by using $K_b = \frac{x^2}{0.1 - x}$. Using the rule of 100, it can be assumed that $0.1 - x \cong 0.1$. This yields $5.6 \times 10^{-11} = \frac{x^2}{0.1}$

$$x = \sqrt{5.6 \times 10^{-12}}$$

$$[OH^-_{(aq)}] = 2.4 \times 10^{-6}\ mol/L$$

$$pOH = -\log[OH^-_{(aq)}] = -\log(2.4 \times 10^{-6}\ mol/L) = 5.62$$

$$pH = 14.00 - 5.62 = 8.38$$

The solution pH is 8.38.

27. Open Response

$$K_{eq} = \frac{[CO][H_2]}{[CO][H_2O]}$$

$$\sqrt{4.6} = \sqrt{\frac{(0.62 + x)^2}{(0.50 - x)^2}}$$

$$x = 0.14$$

$$[H_2] = (0.62 + x) = 0.76\ M$$

28. D

$NaHSO_3$ is acidic in nature, as it contains a replaceable hydrogen atom.

29. B

Since oxalic acid, $HOOCCOOH_{(aq)}$, has two protons to give away, you would expect a titration curve with two sharp rises, presuming both steps are quantitative. This is shown in the graph in alternative B.

30. 0.16

$NaOH_{(aq)} + HCl_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$

n_2	n_1
27.0 mL	18.0 mL
$C = ?$	0.24 mol/L

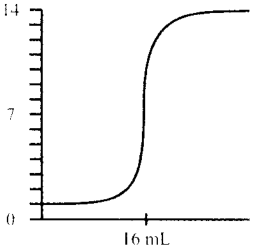
$$n_1 = C \times V = (0.24\ mol/L) \times (18.0\ mL) = 4.3\ mmol$$

$$n_2 = 4.3\ mmol \times \frac{1}{1} = 4.3\ mmol$$

$$[NaOH_{(aq)}] = \frac{n}{V} = \frac{4.3\ mmol}{27.0\ mL}$$

$$= 0.16\ mol/L$$

Part B – Open Response



Bromothymol blue would be a suitable indicator because the midpoint of the rapid rise in the curve, at pH 7, is in the middle of the range of the colour change for bromothymol blue (6.0 to 7.6).

The equivalence point should be marked where the volume is 16 mL and pH is 7.0.

31. Part A – Open Response

Since all reactions are exothermic, a decrease in temperature would cause the equilibrium of each reaction to shift to the right, producing more of the desired products. However, a decrease in temperature also causes a drop in the rate of reaction since molecules lose kinetic energy, which in turn results in fewer effective collisions.

Part B – Open Response

The production of sulfur trioxide, $SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$, is the rate determining step, as a catalyst is added to this step only. The yield of this reaction is increased by decreasing the temperature, but since that slows the rate even further, an increase in pressure is preferred, as this would result in more $SO_{3(g)}$ being produced without slowing the rate.

Part C – Open Response

The production of sulfuric acid is highly exothermic. With so much heat energy being released, it is economically wise to use this energy to generate electricity for the plant.

32. Part A – Open Response

Barium sulfate has a very low K_{sp} and is insoluble in water; therefore, it cannot be absorbed into the blood through the intestines. It is simply passed out without being absorbed. Barium nitrate has a much higher K_{sp} and is much more soluble. It can be absorbed into the blood and possibly cause harm to the body.

Part B – Open Response

$MgSO_{4(aq)} + Ba^{2+}_{(aq)} \rightleftharpoons BaSO_{4(s)} + Mg^{2+}_{(aq)}$

Magnesium sulfate is water soluble, so it is easily absorbed by the body. When it comes in contact with the barium ions, it forms a precipitate, allowing the barium to be safely ejected or excreted from the body.

33. Part A – Open Response

$H^{+}_{(aq)} + HCO^{-}_{3(aq)} \rightleftharpoons H_2CO_{3(aq)} \rightleftharpoons CO_{2(g)} + H_2O_{(l)}$, where HCO_3^- and H_2CO_3 are conjugate acid/base pairs.

Part B – Open Response

Proper enzyme function, cell structure, and the transport of oxygen are all dependent on the pH of blood remaining around 7.4.

Part C – Open Response

Metabolic processes (which increase during exercise) increase the amount of H^+ ions, which in turn shifts the equilibrium to use up these ions and produce more $CO_{2(g)}$. Carbon dioxide must be eliminated from the body or else the pH of the blood will increase too much. The lungs remove excess $CO_{2(g)}$ and the kidneys remove $HCO^{-}_{3(aq)}$ to help shift the equilibrium to the left and thereby decrease blood pH.

Self Test

Open Response

Use the following information to answer the next question.

$HOCl_{(aq)}$, a weak acid, is the active ingredient used in the disinfecting of swimming pools. It can be formed by adding $Ca(OCl)_{2(s)}$ tablets to pool water.

The pH of a swimming pool should be kept between 7.2 and 7.8 so that the equilibrium $[HOCl_{(aq)}]$ is optimal. Phenol red is used by lifeguards to test pH. Based on the test results with phenol red, a lifeguard may

- adjust the pH by adding $Na_2CO_{3(s)}$
- adjust the pH by adding $HCl_{(aq)}$
- not do anything

1. Identify two characteristics of this system, or of any system, at equilibrium.

Use the following information to answer the next question.

The given chemical equation represents the formation of hydrogen iodide.

$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

2. The equilibrium constant for the given equation is

A. $\frac{2[HI]}{[H_2][I_2]}$

B. $\frac{[HI]^2}{[H_2][I_2]}$

C. $\frac{[H_2][I_2]}{[HI]^2}$

D. $\frac{[H_2][I_2]}{2[HI]}$

Use the following information to answer the next question.

The conversion of ice into water is an endothermic process.

Ice + Heat $\rightleftharpoons$ Water

3. Which of the following changes will shift the equilibrium of the process in the backward direction?

A. An increase in the temperature of the system

B. An increase in the pressure of the system

C. A decrease in the pressure of the system

D. An increase in the amount of ice

4. For the conversion of a liquid into a gas, the type of entropy change is represented by

A. $\Delta S = 0$

B. $\Delta S > 0$

C. $\Delta S < 0$

D. $\Delta S \in (0, 1)$

Open Response

5. Consider the following exothermic reaction.

$C_3H_{8(g)} + 5O_{2(g)} \overset{?}{\rightarrow} 3CO_{2(g)} + 4H_2O_{(g)}$

Explain, in terms of increasing or decreasing entropy and enthalpy, whether or not the reaction will reach equilibrium.

____

____

____

____

____

6. The solubility of a solute is **best** determined from which type of solution?

A. A saturated solution

B. Any solution at 25°C

C. An unsaturated solution

D. A supersaturated solution

7. The dissociation constant (K_a) of acids indicates the measure of the relative strength of

A. halogen acids **B.** diprotic acids

C. strong acids **D.** weak acids

Use the following information to answer the next multipart question.

8. Consider the following equilibrium:

$Al(H_2O)_4(OH)^+_{2(aq)}$
$\rightleftharpoons Al(H_2O)_3(OH)_{3(s)} + H^+_{(aq)}$

Part A

Open Response

Some $HCl_{(aq)}$ is added to the equilibrium. What happens to the amount of solid $Al(H_2O)_3(OH)_3$? Explain.

____

____

____

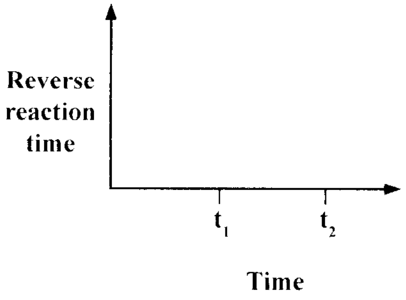
____

____

Part B

Open Response

The HCl is added at time t_1 and equilibrium is re-established at time t_2. On the given axis, sketch what happens to the reverse reaction rate.



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9. Which of the following statements about buffers is **true**?
   - **A.** Buffers resist pH changes to a solution.
   - **B.** Buffers increase the acidity of a solution.
   - **C.** Buffers increase the alkalinity of a solution.
   - **D.** Buffers decrease the alkalinity of a solution.

Open Response

Use the following information to answer the next question.

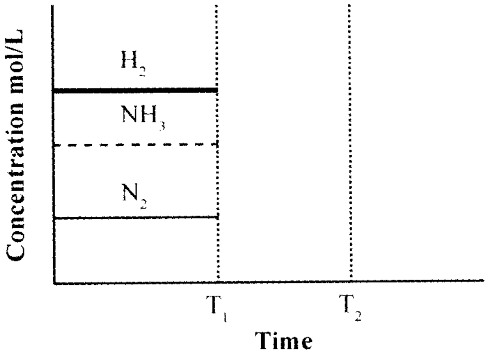
The formation of ammonia gas from its elements occurs according to the following reaction:

$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

10. On the given graph, show what would happen to the equilibrium concentrations of the reactants and products given the following stresses at times T_1 and T_2. Be sure to include a brief explanation for each change in the graph.

At time T_1, hydrogen gas is added to the system.

At time T_2, the volume of the system is increased.



Numerical Response

Use the following information to answer the next question.

Iron(III) thiocyanate is a dark red compound often made from the combination of iron(III) nitrate and potassium thiocyanate. The net reaction is as follows.

$Fe^{3+}_{(aq)} + SCN^{-}_{(aq)} \leftrightharpoons Fe[SCN^{2+}]_{(aq)}$

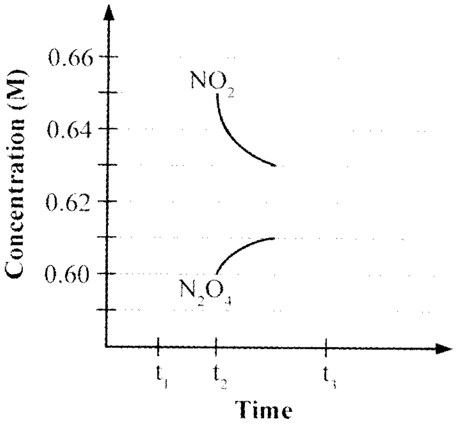
In an experiment to find the K_{eq} of $Fe[SCN^{2+}]$, standard solutions of 0.002 M $Fe(NO_3)_3$ and 0.002 M KSCN were combined at a constant temperature. A table of the experimental results follows. The concentration of $FeSCN^{2+}$ was taken using a spectrophotometer once the reaction ceased changing colour.

	Fe^{3+} +	SCN^{-} ≫	$FeSCN^{2+}$
Initial moles	2.00×10^{-5}	2.00×10^{-5}	0
Change			$+2.5 \times 10^{-6}$
Final moles			
Final volume (L)	0.02	0.02	0.02
Final concentration (M)			1.25×10^{-4}

11. Fill in the missing values in the given data table and find K_{eq} for $Fe[SCN^{2+}]$ ____.

Use the following information to answer the next multipart question.

12. Consider the following diagram for the equilibrium $N_2O_{4(g)} \rightleftharpoons 2NO_{2(g)}$
$\Delta H = +59$ kJ.



Part A

Open Response

Calculate the value of K_{eq} at t_1.

____

Part B

Open Response

Calculate the value of K_{eq} at t_3.

____

13. The solubility of AgI in NaI solution is less than its solubility in pure water because

A. AgI forms a complex with NaI

B. of the common ion effect of iodide ions

C. the shift in equilibrium is toward the right

D. the solubility product of AgI is less than that of NaI

Numerical Response

14. Express the solubility of iron(III) hydroxide in a 0.015 M solution of sodium hydroxide as a number multiplied by 10^{-34}. The solubility product constant for iron(III) hydroxide is 1.6×10^{-39}.

15. Part A

Open Response

Write the net ionic equation for the reaction between $Pb(NO_3)_{2(aq)}$ and $NaCl_{(aq)}$.

Part B

Open Response

Determine, with calculations, whether a precipitate will form when 15.0 mL of 0.500 M $Pb(NO_3)_2$ is added to 35.0 mL of 0.085 M NaCl.

Numerical Response

Use the following information to answer the next question.

A 0.500 mol/L solution of hydrazine ($N_2H_{4(aq)}$) contains the following equilibrium concentrations.

$$[N_2H_{4(aq)}] = 0.498\ \text{mol/L}$$
$$[OH^-_{(aq)}] = 2.14 \times 10^{-3}\text{mol/L}$$
$$[N_2H^+_{5(aq)}] = 2.14 \times 10^{-3}\ \text{mol/L}$$

16. The K_b for hydrazine, in scientific notation, is $a.bc \times 10^{-d}$. The values of a, b, c, and d are, respectively, ____, ____, ____, and ____.

Open Response

17. Using calculations, show why the electrical conductivity of 1.0 M H_2CO_3 will be less than that for 0.10 M HCl.

18. Which of the following salt solutions is neutral?

A. NaCl solution

B. NH_4Cl solution

C. $NaHCO_3$ solution

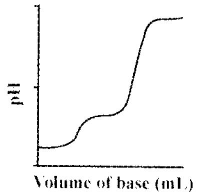
D. $Na_2C_2O_4$ solution

Numerical Response

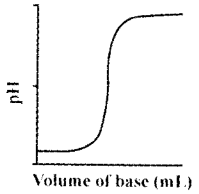
Use the following information to answer the next question.

When equally concentrated solutions of $HNO_{3(aq)}$, $CH_3COOH_{(aq)}$, $HOOCCOOH_{(aq)}$, and $Na_2S_{(aq)}$ were titrated with either a strong acid or a strong base, the following titration curves were obtained:

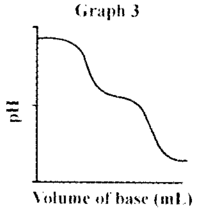
Graph 1



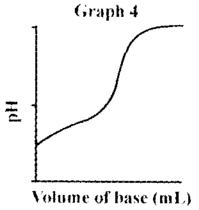
Graph 2



Graph 3



Graph 4



19. Match each of the graphs, as numbered above, with the corresponding titration species listed below.

$HNO_{3(aq)}$ ____

$CH_3COOH_{(aq)}$ ____

$HOOCCOOH_{(aq)}$ ____

$Na_2S_{(aq)}$ ____

Open Response

20. A 100 mL sample of 0.25 M acetic acid, $CH_3COOH_{(aq)}$, is titrated with 0.25 M of sodium hydroxide. What is the pH at the equivalence point?

Note: K_a for acetic acid is 1.8×10^{-5}.

SOLUTIONS

1. **OR**
2. **B**
3. **C**
4. **B**
5. **OR**
6. **A**
7. **D**
8. Part A- **OR**
   Part B- **OR**
9. **A**
10. **OR**
11. **163**
12. Part A- **OR**
    Part B- **OR**
13. **B**
14. **4.7**
15. Part A- **OR**
    Part B- **OR**
16. **9,2,0,6**
17. **OR**
18. **A**
19. **2 4 1 3**
20. **OR**

1. Open Response

Equilibrium systems have no measurable macroscopic changes in system properties (e.g., pH, temperature, concentration, mass, amount of reactants or products, colour, and pressure all remain constant).

and/or

Equilibrium systems have dynamic microscopic properties (e.g., rate of forward reaction equals rate of reverse reaction).

and/or

Equilibrium systems are closed systems (e.g., no energy or matter can enter or leave the system).

2. B

On applying the law of mass action on the given reaction,

Rate of forward reaction $\propto [H_2][I_2] = K_f[H_2][I_2]$

Rate of backward reaction $\propto [HI]^2 = K_b[HI]^2$

At equilibrium, rate of forward reaction = rate of backward reaction.

$$K_f[H_2][I_2] = K_b[HI]^2$$

$$\frac{[HI]^2}{[H_2][I_2]} = \frac{K_f}{K_b} = K = \text{equilibrium constant}$$

Therefore, $K = \dfrac{[HI]^2}{[H_2][I_2]}$.

3. C

The conversion of ice into water involves a decrease in volume, i.e., when ice melts into water, its volume decreases. According to Le Châtelier's principle, with the decrease in pressure, the volume of the system decreases and the reaction moves in such a direction that the volume increases, i.e., in the backward direction.

Since the reaction is endothermic, increasing the temperature of the system will favour the endothermic process, driving the reaction in the forward direction. Similarly, increasing the amount of ice, which is on the side of the reactants, along with heat, will drive the reaction in the forward direction. In either case, the equilibrium is altered in such a way as to counteract the additional stress, which in this case is adding something else to the reactants side.

4. B

Entropy is a measure of the randomness or the disorder of a system. When a liquid is converted into a gas, the randomness of the system increases, and thus the entropy change for the system is positive.

5. Open Response

Entropy increases in the forward reaction.

Enthalpy decreases forward reaction.

Since both favour products, equilibrium will not be attained, or the reaction will go to completion.

6. A

The maximum possible concentration of a solute in a solution under normal circumstances defines the solubility of that solute. A solution containing the maximum concentration of solute normally soluble at a given temperature is a saturated solution.

7. D

The dissociation of a weak acid in water can be represented as

$$HA + \text{water} \rightleftharpoons A^-_{(aq)} + H^+_{(aq)}$$

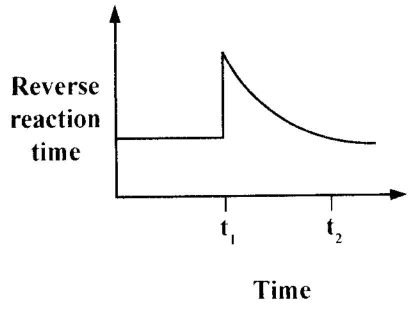
$K_a = \dfrac{[A^-][H^+]}{[HA]}$, where K_a is the dissociation constant of the weak acid.

The greater the value of K_a, the greater the dissociation of the acid. Thus, K_a indicates the measure of the relative strength of weak acids.

8. Part A – Open Response

The amount of solid decreases because the equilibrium shifts left.

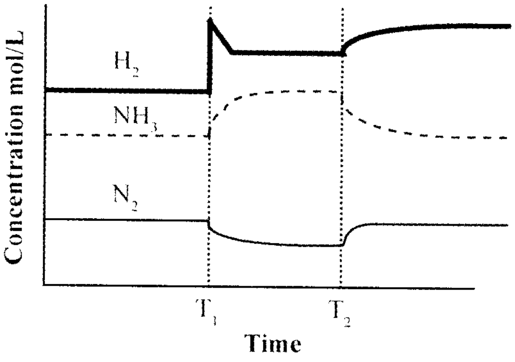
Part B – Open Response



9. A

Buffers resist pH changes to a solution.

10. Open Response



At time T_1, the $[H_{2(g)}]$ sharply increases. Then, as the equilibrium shifts to the right to use up the excess $H_{2(g)}$, the $[NH_{3(g)}]$ increases and $[H_{2(g)}]$ and $[N_{2(g)}]$ decrease. At time T_2, the volume of the system is increased, which means the pressure decreases. The equilibrium will shift to the left to compensate, creating more gas molecules, thereby increasing the pressure. The $[H_{2(g)}]$ and $[N_{2(g)}]$ increase, while $[NH_{3(g)}]$ decreases.

11. 163

	Fe^{3+} +	SCN^- ≫	$FeSCN^{2+}$
Initial moles	2.00×10^{-5}	2.00×10^{-5}	0
Change	-2.5×10^{-6}	-2.5×10^{-6}	$+2.5 \times 10^{-6}$
Final moles	1.75×10^{-5}	1.75×10^{-5}	2.5×10^{-6}
Final volume (L)	0.02	0.02	0.02
Final concentration (M)	8.75×10^{-4}	8.75×10^{-4}	1.25×10^{-4}

$$K_{eq} = \frac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^-]}$$

$$= \frac{1.25 \times 10^{-4}}{(8.75 \times 10^{-4})^2} = 163$$

12. Part A – Open Response

$$K_{eq} = \frac{[NO_2]^2}{[N_2O_4]} = \frac{(0.65)^2}{0.60} = 0.70$$

Part B – Open Response

$$K_{eq} = \frac{(0.63)^2}{0.61} = 0.65$$

13. B

$AgI_{(s)} \rightleftharpoons Ag^+_{(aq)} + I^-_{(aq)}$
$K_{sp} = [Ag^+][I^-]$

NaI completely ionizes in a solution to give Na^+ and I^- ions. When AgI is added to an NaI solution, the concentration of iodide ions increases and the equilibrium shifts in the backward direction because of the common ion effect. The shift in the equilibrium suppresses the solubility of AgI in an NaI solution.
No common ions are present in water. Hence, the solubility of AgI is comparatively higher in water.
No reaction takes place between NaI and AgI.
An increase in the concentration of iodide ions will shift the equilibrium toward the left-hand side.
The solubility product of AgI is comparatively less than that of NaI. While AgI exhibits low solubility, NaI is quite soluble in water.

14. 4.7

$Fe(OH)_{3(s)}$	$\rightleftharpoons$	Fe^{3+}	+	$3OH^-_{(aq)}$
x		x		$3x + 0.015$

$K_{sp} = [Fe^{3+}][OH^-]^3 = x \times (3x + 0.015)^3$
Assume $3x + 0.015 \cong 0.015$.
Check assumption: $\frac{0.015}{1.6 \times 10^{-39}} > 100$, thus the assumption is correct.

$1.6 \times 10^{-39} = x \times (0.015)^3 x = \frac{1.6 \times 10^{-39}}{3.375 \times 10^{-6}}$

$x = 4.7 \times 10^{-34}$
Thus, because $Fe(OH)_{3(s)}$ and x are in a 1:1 ratio, the molar solubility of $Fe(OH)_{3(s)}$ is 4.7×10^{-34}.

15. Part A – Open Response

$Pb^{2+}_{(aq)} + 2Cl^-_{(aq)} \rightarrow PbCl_{2(s)}$

Part B – Open Response

$[Pb^{2+}] = 0.050 \text{ M} \times \frac{15.0 \text{ mL}}{50.0 \text{ mL}}$
$= 0.015 \text{ M}$
$[Cl^-] = 0.085 \text{ M} \times \frac{35.0 \text{ mL}}{50.0 \text{ mL}}$
$= 0.0595 \text{ M}$
Trial $K_{sp} = [Pb^{2+}][Cl^-]^2$
$= (0.015)(0.0595)^2$
$= 5.3 \times 10^{-5}$
K_{sp} for $PbCl_2 = 1.2 \times 10^{-5}$. Since Trial $K_{sp} > K_{sp}$, a precipitate does form.

16. 9,2,0,6

Assume that the hydrazine solution is basic because its $\left[OH^-_{(aq)}\right]$ is greater than $1.0 \times 10^{-7} mol/L$ by at least four orders of magnitude (a factor of at least 10^4). So, a basic equilibrium is written as

$$N_2H_{4(aq)} + H_2O_{(l)} \rightleftharpoons N_2H^+_{5(aq)} + OH^-_{(aq)}.$$

Therefore, $K_b = \dfrac{\left[N_2H^+_{5(aq)}\right]\left[OH^-_{(aq)}\right]}{\left[N_2H_{4(aq)}\right]}$.

Substituting the concentrations provided gives you

$$K_b = \frac{(2.14 \times 10^{-3}\ mol/L)(2.14 \times 10^{-3}\ mol/L)}{(0.498\ mol/L)}$$

and $K_b = 9.20 \times 10^{-6}\ mol/L$.

In scientific notation $(a.bc \times 10^{-d})$, a, b, c, and d are, respectively, 9, 2, 0, and 6.

17. Open Response

For 0.10 M HCl, a strong acid,

$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)}Cl^-_{(aq)}$.

Total concentration = 0.10 M + 0.10 M = 0.20 M

For 1.0 M H_2CO_3, a weak acid,

	$H_2CO_{3(aq)} + H_2O_{(l)}$	$\rightleftarrows$	$H_2CO^-_{3(aq)}$	$+H_3O^+_{(aq)}$
[I]	1.0		0	0
[C]	$-x$		$+x$	$+x$
[E]	$1.0 - x$		x	x

$$K_a = \frac{\left[H_3O^+\right]\left[HCO_3^-\right]}{\left[H_2CO_3\right]} = 4.3 \times 10^{-7}$$

$$\frac{(x)(x)}{1.0 - x} = 4.3 \times 10^{-7}$$

$x = 6. \times 10^{-4} M = \left[H_3O^+\right] = \left[HCO_3^-\right]$

Total ion concentration

$= 6.6 \times 10^{-4}\ M + 6.6 \times 10^{-4}\ M = 1.3 \times 10^{-3}\ M$

Therefore, the smaller the ion concentration, the lower the conductivity.

18. A

NaCl is a salt of a strong base, NaOH, and a strong acid, HCl. The salt dissociates completely in an aqueous solution to produce the strong acid and the strong base in equal proportions. Therefore, the solution is neutral.

19. 2 4 1 3

The number of equivalence points (steep drops or jumps) in a pH titration curve corresponds to the number of protons quantitatively added to or removed from the sample acid or base.

The pH of an equivalence point is contingent largely on the acidity/basicity of the conjugate produced on neutralization. If a weak acid conjugate is produced on neutralization, the equivalence point pH lies below 7. If a weak base conjugate is produced, the equivalence point pH lies above 7. If $H_2O_{(l)}$ is the conjugate, the equivalence point pH is 7.

When $HNO_{3(aq)}$ is titrated with a strong base, the Brønsted–Lowry net reaction is

$H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2H_2O_{(l)}$, and the equivalence point pH is 7, as there is only one equivalence point.

This corresponds to Graph 2—of course, the graph will start at low pH and then rise because a base titrant is being added to an acid sample.

When $CH_3COOH_{(aq)}$ is titrated, the Brønsted–Lowry net reaction is $CH_3COOH_{(aq)} + OH^-_{(aq)}$

$\rightarrow CH_3COO^-_{(aq)} + H_2O_{(l)}$.

The graph will start at a lower pH and then rise with one equivalence point. The equivalence point pH ought to be above 7, since the $CH_3COO^-_{(aq)}$ produced is a weak base —this is clearly a description of Graph 4.

When $HOOCCOOH_{(aq)}$ is titrated, the Brønsted–Lowry net reactions are

i. $HOOCCOOH_{(aq)} + OH^-_{(aq)}$

$\rightarrow HOOCCOO^-_{(aq)} + H_2O_{(l)}$

ii. $HOOCCOO^-_{(aq)} + OH^-_{(aq)}$

$\rightarrow OOCCOO^{2-}_{(aq)} + H_2O_{(l)}$

The graph of this titration will start at a lower pH and rise through two equivalence points. The first equivalence point will have a pH below 7 (the conjugate is the acidic amphiprotic, $HOOCCOO^-_{(aq)}$).

The second equivalence point will have a pH above 7 (the conjugate is the weak base, $OOCCOO^{2-}_{(aq)}$).

This corresponds to Graph 1.

By default, the titration curve for $Na_2S_{(aq)}$ (really $S^{2-}_{(aq)}$, a weak base) must be Graph 3. It is the only graph that starts high and then drops. (Note: Amphiprotics tend to form acidic solutions in water if they lie above $HCO_{3(aq)}$ in the Relative Strengths of Acids and Bases Table and basic solutions in water if they lie below $HCO^-_{3(aq)}$ in the table. $HCO^-_{3(aq)}$ itself forms basic aqueous solutions.)

20. Open Response

$NaOH_{(aq)} + CH_3COOH_{(aq)} \rightleftharpoons NaCH_3COO_{(aq)} + H_2O_{(l)}$

From the given equation, you can see that at the equivalence point there is only sodium acetate ions. Because acetic acid and sodium hydroxide are the same concentration and they react 1:1, the solution volume at the equivalence point is 200 mL.

$$C \text{ of } NaCH_3COOH = \frac{0.25\ M \times 0.1\ L}{0.2\ L} = 0.125\ M$$

Thus, $c_{CH_3COO^-} = 0.125\ M$

140296-table1.eps

$CH_3COO_{(aq)}$	$H_2O_{(l)}$	$OH^-_{(aq)}$	$CH_3COO_{(aq)}$
0.125 M		0	0
$-x$		$+x$	$+x$
$0.125 - x$		x	x

The acetate ions act as a base. Since $K_a \times K_b = K_w$, you can find K_b.

$$K_b = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.5 \times 10^{-10}$$

$$K_b = 5.5 \times 10^{-10} = \frac{[OH^-][CH_3COOH]}{[CH_3COO^-]} = \frac{x^2}{0.125 - x}$$

Assume $0.125 - x \cong 0.125$. Check: $\frac{0.125}{5.5 \times 10^{-10}} > 100$, thus the assumption is valid.

$$5.5 \times 10^{-10} = \frac{x^2}{0.125}$$

$$x = \sqrt{6.875 \times 10^{-11}} = 8.29 \times 10^{-6}\ M = [OH^-]$$

$$pOH = -\log[OH^-] = 5.08$$

$$pH = 14 - pOH = 8.9$$

Electrochemistry

Electrochemistry

Table of Correlations

Specific Expectation	Practice Questions	Self Test Questions
12.4.1 Understanding Basic Concepts		
12.4.1.1 *demonstrate an understanding of oxidation and reduction in terms of the loss and the gain of electrons or change in oxidation number*	1	1
12.4.1.2 *identify and describe the functioning of the components in galvanic and electrolytic cells*	2	2
12.4.1.3 *describe electrochemical cells in terms of oxidation and reduction half-cells whose voltages can be used to determine overall cell potential*	3, 4	3
12.4.1.4 *describe the function of the hydrogen half-cell as a reference in assigning reduction potential values*	5, 6	4
12.4.1.5 *demonstrate an understanding of the interrelationship of time, current, and the amount of substance produced or consumed in an electrolytic process (Faraday's law)*	7, 8	5
12.4.1.6 *explain corrosion as an electrochemical process, and describe corrosion-inhibiting techniques*	10	6
12.4.2 Developing Skills of Inquiry and Communication		
12.4.2.1 *use appropriate scientific vocabulary to communicate ideas related to electrochemistry*	11	
12.4.2.2 *demonstrate oxidation-reduction reactions through experiments, and analyse these reactions*	12, 13, 14	7, 8
12.4.2.3 *write balanced chemical equations for oxidation-reduction systems, including half-cell reactions*	15, 16	9
12.4.2.4 *determine oxidation and reduction half-cell reactions, direction of current flow, electrode polarity, cell potential, and ion movement in typical galvanic and electrolytic cells, including those assembled in the laboratory*	17, 18, 19	10
12.4.2.5 *predict the spontaneity of redox reactions and overall cell potentials by studying a table of half-cell reduction potentials*	20, 21	11
12.4.2.6 *solve problems based on Faraday's law*	9	12
12.4.2.7 *measure through experimentation the mass of metal deposited by electroplating, and apply Faraday's law to relate the mass of metal deposited to the amount of charge passed*	22, 23, 24	13
12.4.3 Relating Science to Technology, Society, and the Environment		
12.4.3.1 *describe examples of common galvanic cells and evaluate their environmental and social impact*	25	
12.4.3.2 *explain how electrolytic processes are involved in industrial processes*	26	
12.4.3.3 *research and assess environmental, health, and safety issues involving electrochemistry*	27	

12.4.1.1 *demonstrate an understanding of oxidation and reduction in terms of the loss and the gain of electrons or change in oxidation number*

OXIDATION AND REDUCTION

INTRODUCTION

The history of oxidation and reduction chemistry comes mostly from metallurgy. In ancient times, people realized that useful pure metals could be extracted from ores through a variety of processes. The term used to describe these processes was called reduction. The ores were combined with a reducing agent and then heated to produce metals such as copper, tin, and iron. The most commonly used reducing agents were carbon monoxide, hydrogen gas, and carbon. It was later understood that the content of the ores was the pure form of the metal combined with oxygen. These metals were oxidized much in the same way as corrosion and combustion involve a chemical reaction with oxygen. It was later discovered that oxygen was not the only substance that could cause reactions that appeared similar to oxidation reactions. The term oxidation was thus broadened to include numerous types of combustion and corrosion reactions. Below is an example of the oxidation of a pure copper by oxygen to an oxide followed by its reduction back to the pure metal by a hydrogen gas reducing agent. In the first reaction, oxygen gas acts as the oxidizing agent.

Example

$2Cu_{(s)} + O_{2(g)} \rightarrow 2CuO_{(s)}$

$CuO_{(s)} + H_{2(g)} \rightarrow Cu_{(s)} + H_2O_{(l)}$

ELECTRON TRANSFER AND REDOX CHEMISTRY

In order to better understand what is occurring in oxidation and reduction reactions, it is important to understand that what is actually occurring is a transfer of electrons. Oxidation is a process in which electrons are lost and reduction is a process in which electrons are gained. An easy way to remember this is that when electrons are gained, the charge is **reduced**. In oxidation-reduction reactions or **redox** reactions, both processes occur, so that the electrons lost by the oxidation are gained during the reduction. Because of this, the electrons do not appear in the combined reaction. In order to write chemical reactions with the electrons included, the redox reaction must be split into half-reactions with one half reaction being the oxidation reaction and the other the reduction. It is important to note that the number of electrons lost and gained must be equal, so the half reactions may need to be multiplied by integer values. The following example is a simple single replacement redox reaction where iron is oxidized and **copper (II) sulfate** is reduced. The overall reaction is split up into the two half reactions representing the oxidation and reduction.

Example

$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$

$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

OXIDATION NUMBER

The above treatment of redox theory works well for atoms and monatomic ions, but is not sufficient to describe the redox reactions that occur with more complex molecules such as metal oxides. Therefore, it is necessary to have a way to know what atoms in the more complex molecules are undergoing oxidation or reduction and by how many electrons. Oxidation numbers provide a method to keep track of these changes. An oxidation number is an arbitrary positive or negative number that describes the redox state of an atom in a molecule. The number is based on the assumption that the more electronegative atom in a covalent bond will have all of the electrons from that bond and therefore have a negative number. It is important to understand that this number does not represent an actual charge and is just a system of bookkeeping. If the oxidation number of an atom on the reactant side of a chemical equation is different on the product side, a redox reaction has occurred. The oxidation numbers of all atoms in a molecule must add up to the charge of the molecule. This allows the oxidation numbers of all atoms in a molecule to be determined if the oxidation number of certain atoms is known. The following table gives the most common oxidation numbers of some elements when they are in molecules:

Atom or Ion	**Oxidation Number**
all atoms in elements	0
hydrogen	+1
hydrogen in hydrides	−1
oxygen	−2
oxygen in peroxides	−1
all monatomic ions	charge on ion

Practice

1. In which of the following molecules does carbon have a negative oxidation number?

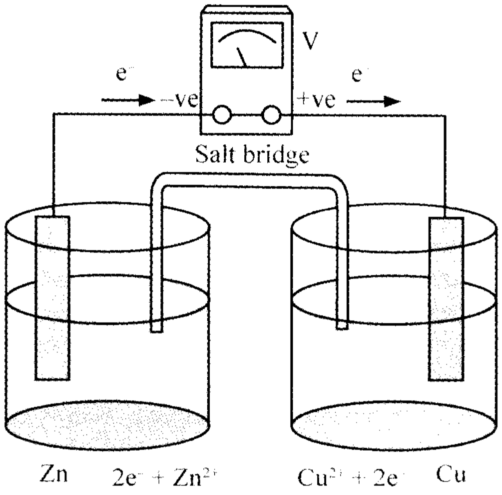
   A. CH_4 **B.** $HCCl_3$

   C. CO_2 **D.** CO

12.4.1.2 *identify and describe the functioning of the components in galvanic and electrolytic cells*

GALVANIC CELLS

Galvanic or voltaic cells provide a way to study electrochemical reactions and can also be used to supply power. The electrons that are normally transferred from atom to atom in a redox reaction are instead passed through an external circuit which creates a current. This is accomplished by keeping the oxidizing and reducing agents separated. Below is a depiction of a typical galvanic cell used in scientific research:

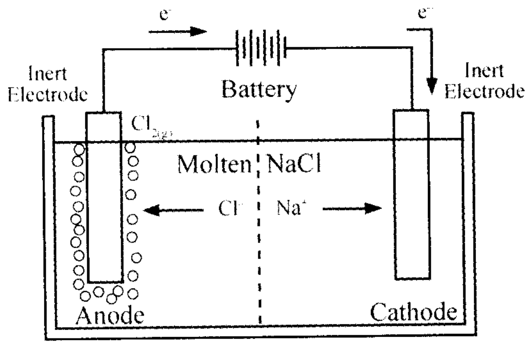


Zinc-copper galvanic cell

The galvanic cell consists of three main parts: two electrodes and an electrolyte. In this case, the two separate compartments represent the electrodes, with oxidation occurring on the zinc side and reduction occurring on the copper side. The site of the reduction reaction is called the cathode and the site of the oxidation reaction is called the anode, which means electrons flow from the anode to the cathode through a wire that can be connected to a voltmeter. The electrons are generated at the anode as the zinc metal is oxidized to Zn^{2+}. At the cathode, they combine with Cu^{2+} ions to form solid copper. Therefore, the electrode in the anode dissolves while the electrode in the cathode increases in mass. This occurs spontaneously because of the difference in electrochemical potential of the two metals. The electrolyte in the two electrodes acts to balance the charge by providing unreactive negative ions on the anode side and positive ions on the cathode side. The salt bridge provides a means for this transfer to counteract the charge difference that develops as the electrons move. This charge difference would prevent further electron flow if the salt bridge was absent.

ELECTROLYTIC CELLS

Electrolytic cells differ from galvanic cell in that they need an external power source to initiate redox chemistry. The reason for this is that the combination of potential oxidizing and reducing agents do not have a cell potential that results in a spontaneous reaction. The figure below is a typical electrolytic cell consisting of two inert carbon electrodes immersed in molten sodium chloride:



Electrolysis of NaCl

Typical design for electrolytic cell

In this cell, only Cl^- and Na^+ are available to do redox chemistry as the electrodes themselves are inert. Therefore, the external battery is required to force electrons in one direction and establish a cathode and anode. Again, reduction occurs at the cathode because that is the electrode where electrons are available for reduction of Na^+ to Na. At the anode, oxidation occurs as the battery provides the electrochemical force to remove electrons from the Cl^-. The purpose of this type of cell is most often to produce a desired product, in this case, sodium and chlorine gas.

Practice

2. Which of the following is the major difference between a galvanic and electrolytic cell?

A. The anode

B. The cathode

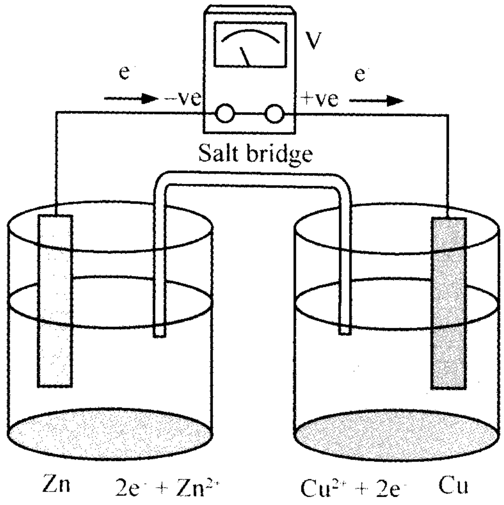
C. Spontaneity of the redox reaction

D. Oxidizing and reducing agents

12.4.1.3 *describe electrochemical cells in terms of oxidation and reduction half-cells whose voltages can be used to determine overall cell potential*

HALF-CELLS

When describing the redox reactions that are occurring in electrochemical cells, a useful notation exists that summarizes the important information quickly. This "shorthand" notation is called cell notation and relies on the fact that electrochemical cells can be broken up into two half-cells consisting of an electrode and electrolyte which represent the two half reactions occurring in the redox reaction. Consider the figure of the copper-zinc galvanic cell below.



One half-cell has copper metal and a solution of copper ions while the other has zinc metal and a solution of zinc ions. In this particular case, the electrolyte is potassium chloride so the zinc and copper ions can be put into solution as $CuCl_2$ and $ZnCl_2$. To write the cell, the information is given as $Zn_{(s)} | Zn(Cl)_{2(aq)} || Cu(Cl)_{2(aq)} | Cu_{(s)}$ where the single line indicates a phase boundary such as liquid and solid and the double line represents a physical boundary, like the salt bridge between the half-cells. Cell notation presents the anode half-cell reaction on the left and the cathode half-cell reaction on the right. Once the design of the cell is considered, it can be determined which half reactions occur. The reduction half-reaction occurs at the cathode and has the form $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ and the oxidation half-reaction occurs at the anode and has the form $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$. Cell notation can also be used for electrolytic cells. For example, a cell that consists of two inert carbon electrodes immersed in a solution of potassium iodide will have the cell notation $C_{(s)} | K^+_{(aq)}, I^-_{(aq)} | C_{(s)}$. Note that there is no double line because of the lack of a physical barrier in this cell. The cell notation represents what is present at the creation of the cell, not what is produced when it is active.

CELL POTENTIAL

In the figure above, the two half-cells are connected together by a conducting wire that is attached to a voltmeter. When the electrons from the redox reaction move through the wire, they cause a current which corresponds to a voltage in the metre. This voltage is used as a measure of the "cell potential". The cell potential depends on which half-cells are used and is a direct measure of how good of a reducing agent is at the anode and how good of an oxidizing agent is at the cathode. The cell potential is also dependent on the concentrations of ions in the electrolyte. Each half-cell can be though of as having a certain amount of "oxidation potential" or "reduction potential". For a given half reaction, the potential of the oxidation reaction has the opposite sign of the reduction reaction or $E_{red} = -E_{ox}$. The total cell potential is simply the addition of the oxidation potential at that anode and the reduction potential at the cathode. To simplify things, all half-reaction potentials are thought of as reduction potentials. This means the equation to determine cell potential is $E = E_{\text{cathode}} - E_{\text{anode}}$ where E_{cathode} and E_{anode} are the respective reduction potential of the half-reactions occurring at the cathode and anode. It can be seen that in order to have a positive cell potential (spontaneous reaction) the reduction potential of the cathode must be greater than the reduction potential of the anode. This makes sense because reduction occurs at the cathode and the half-reaction with the higher reduction potential will be the one that undergoes reduction. In electrolytic cells, the cell potential can be determined by the amount of voltage that needs to be applied from an external power source in order to initiate the reaction. Half-cell potentials have no real physical meaning as no potential can be established without a redox partner. Connecting various types of half-cells to one another and measuring the voltage is the only way to compare half-cell potentials.

Practice

3. If a hypothetical galvanic cell with a positive cell potential consists of one half-cell that undergoes the reaction $A_{(s)} \rightarrow A^{2+}_{(aq)} + 2e^-$, which of the following hypothetical half-cells could be the other portion of the complete galvanic cell?

A. $B_{(s)} \rightarrow B^{2+}_{(aq)} + 2e^-$

B. $A^{2+}_{(aq)} + 2e^- \rightarrow A_{(s)}$

C. $B^{+}_{(aq)} + e^- \rightarrow B_{(s)}$

D. None of the above

Numerical Response

4. If the cell notation for a galvanic cell is $Ag_{(s)} | AgCl_{(aq)} || AuCl_{3(aq)} | Au_{(s)}$, how many electrons does the redox reaction involve? ____

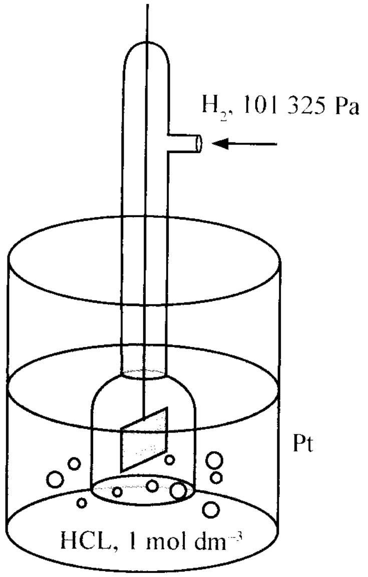
12.4.1.4 *describe the function of the hydrogen half-cell as a reference in assigning reduction potential values*

STANDARD CELLS

They way in which a cell is constructed affects its potential. The concentrations of ions in the electrolyte and the pressure of any gases used in electrodes can affect the potential along with the temperature at which the experiment is conducted. This makes it difficult to compare results from one laboratory to another. Because of this, specific concentrations and conditions are set forth as standards by which all cells can be compared. The standard cell potential ($\Delta E°$) is measured at STP with the concentration of cell components at 1.0 M. This insures that the cell potential of any combination of half-cells will match that same combination in any other laboratory. There is also the standard reduction potential ($E_r°$) for half-cells, which is a measure of the potential of a half cell to attract electrons and by doing so act as the cathode. The attracted electrons act in the reduction half-reaction. The way in which cell potential relates to the reduction potential of the cathode and anode is $\Delta E° = E°_{r(\text{cathode})} - E°_{r(\text{anode})}$. The standard reduction potential of half-reactions cannot be directly measured because it relies on a redox partner, so only the standard cell potential can be measured. However, if every half-cell was measured under standard conditions with reference to a half-cell that was said to have a reduction potential of zero, comparisons could be made. This half-cell with zero reduction potential is the hydrogen half cell.

HYDROGEN HALF-CELL

The standard hydrogen half-cell is an inert platinum electrode in a 1.0 M solution of hydrogen ions and hydrogen gas at standard pressure above the electrode. A figure of the electrode is shown below:



This half cell represents the hydrogen ion reduction half-reaction $2H^+_{(aq)} + 2e^- \rightleftharpoons H_{2(g)}$ that is said to have a standard reduction potential of zero. All other half-reactions can be constructed into half-cells at standard conditions and attached to the hydrogen half-cell. When the total cell potential is measured, if it has a positive value, it means that the oxidizing agent is stronger than H^+. If it is negative, it means the opposite. The positive and negative values only have meaning because when the voltmeter is hooked up, a cathode and anode must be chosen. If hydrogen is always chosen as the anode, the positive reading means that this assumption was correct and hydrogen is oxidized. Negative values can be thought of as incorrectly hooking up the voltmeter because the cathode and anode are actually reversed (H^+ was the stronger oxidizing agent). By this method, standard reduction potential tables can be constructed that are useful for predicting the outcome of redox reactions and electrochemical cell behaviour. In the table, the reactions are all written as reduction reactions and a positive reduction potential means that the substance is a stronger oxidizing agent then H^+ and will act as the cathode in a galvanic cell with the standard hydrogen electrode. Negative values mean that hydrogen will act as the cathode and the oxidization reaction (the reverse of the one on the table) will occur at that anode.

Practice

5. If an electrochemical cell consists of the standard hydrogen electrode (hooked up as anode) and a chromium electrode (hooked up as cathode) and the voltmeter reads −0.74 V, which of the following best describes the system?
   - **A.** Cathode and anode are reversed, oxidation occurs at the hydrogen electrode
   - **B.** Cathode and anode are reversed, reduction occurs at the chromium electrode
   - **C.** Cathode and anode are correctly wired, oxidation occurs at the hydrogen electrode
   - **D.** Cathode and anode are reversed, reduction occurs at the hydrogen electrode

Open Response

6. Write the cell notation and calculate the cell potential for a standard iron-lead cell.

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12.4.1.5 *demonstrate an understanding of the interrelationship of time, current, and the amount of substance produced or consumed in an electrolytic process (Faraday's law)*

12.4.2.6 *solve problems based on Faraday's law*

FARADAY'S LAW

Because electrolysis involves nonspontaneous reactions, electrolytic cells are not useful for current generation. However, they are useful for producing useful chemical species. Most of the time, this involves reactants that are commonly available and by using an external current to force the nonspontaneous reaction, more desirable products can be made. In order to understand electrolysis, it is important to understand how the amount of current applied to the cell relates to the amount of product produced in a given amount of time. When looking at a balanced redox reaction, it is possible to determine how many moles of electrons are necessary to produce one mole of product by using stoichiometry. This relationship between the amount of substance produced and the amount of current applied is called Faraday's law. To understand Faraday's law, it is necessary to understand how current relates to number of electrons. The units of current are amperes or charge per second. Charge is measured in coulombs with the charge of an electron being 1.602×10^{-19} C.

Therefore one mole of electrons has a charge of

$$\frac{1.602 \times 10^{-19}\ \text{C}}{1e^-} \times \frac{6.022 \times 10^{23} e^-}{1\ \text{mol}}$$

$= 9.647 \times 10^4$ C / mol.

This gives the relationship between charge and number of moles that is essential in relating stoichiometry to electricity. To find the amount of charge when a current is applied for a period of time, use the following relationship:
charge(C) = current(A) × time(s). The following example deals with the electrolysis of aluminium:

Example

How many grams of aluminium are produced when a current of 300 mA is passed through a solution of molten aluminium chloride for 1.00 h?

First, it is necessary to figure out the amount of charge produced from the current and time.

$1.00\text{ h} \times \frac{3\,600\text{ s}}{\text{h}} = 3.60 \times 10^{3}\text{ s}$ and

$3.60 \times 10^{3}\text{ s} \times 3.00 \times 10^{-1}\text{ C/s} = 1.08 \times 10^{3}\text{ C}$

now convert this to moles of electrons

$1.08 \times 10^{3}\text{ C} \times \frac{\text{mol}}{9.647 \times 10^{4}\text{ C}} = 1.12 \times 10^{-2}\text{ mol.}$

Because $Al^{3+}_{(aq)} + 3e^{-} \rightarrow Al_{(s)}$, three moles of electrons are required for each mole of solid aluminium.

$\frac{1\text{ mol }Al_{(s)}}{3\text{ mol }e^{-}} \times 1.12 \times 10^{-2}\text{ mol}$

$= 3.73 \times 10^{-3}\text{ mol of }Al_{(s)}$

and finally convert to grams

$3.73 \times 10^{-3}\text{ mol }Al_{(s)} \times \frac{26.98\text{ g}}{\text{mol}} = 0.101\text{ g.}$

Practice

7. Which of the following is the definition of 1 F (one Faraday)?

A. Charge of one mole of electrons

B. Current generated by one mole of electrons

C. Voltage associated with one mole of electrons

D. Number of electrons in one coulomb

Open Response

8. Derive an expression relating the volume of a gas evolved, given a particular current and time in an electrolysis experiment. (assume each mole of electrons gives a mole of gas)

9. Which one of the following aqueous metal nitrate solutions would produce the largest number of moles of solid product if all were subject to the same amount of current for the same time?

A. Silver nitrate

B. Iron (II) nitrate

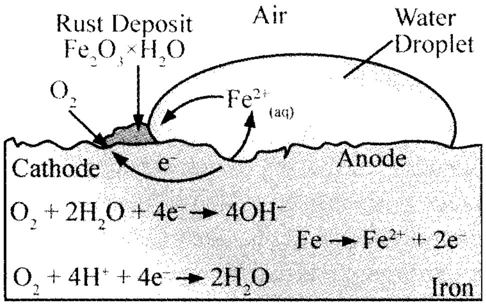
C. Aluminium nitrate

D. Copper (II) nitrate

12.4.1.6 *explain corrosion as an electrochemical process, and describe corrosion-inhibiting techniques*

CORROSION

Corrosion is the spontaneous redox reaction of materials with substances within their environment. Corrosion is most apparent in human society with respect to the oxidation of metals. Because Earth's atmosphere is composed of a large portion of oxygen (a strong oxidant), many metals are easily oxidized through exposure of reduced surfaces to air. The rusting of iron is one of the most relevant examples of corrosion. Since corrosion is a naturally occurring spontaneous redox reaction, it is convenient to talk about rusting in terms of galvanic cells. Rust is a hydrated iron (III) oxide and its formation is a complex process. Rusting occurs on the surface of iron, which acts as the anode, and requires water, an inert cathode (i.e. carbon), and an electrolyte (usually carbonic acid from rain water). The figure below diagrams the rusting process:



The main half reactions occurring in rust formation are $Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$ at the anode and $O_{2(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^-_{(aq)}$ at the cathode. Since this naturally occurring galvanic cell does not have any barriers to product ions reacting, the following reaction occurs: $Fe^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Fe(OH)_{2(s)}$ which can further oxidize by reaction with air to iron (III) hydroxide: $4Fe(OH)_{2(s)} + O_{2(g)} + 2H_2O_{(l)} \rightarrow 4Fe(OH)_{3(s)}$. Finally, the iron (III) hydroxide breaks down to form rust $2Fe(OH)_{3(s)} \rightarrow Fe_2O_3{\cdot}3H_2O_{(s)}$.

CORROSION PREVENTION

SURFACE PROTECTION

The simplest way to protect an iron object from corrosion is to prevent water and oxygen from accessing the metal. This can be done with inexpensive surface coatings such as paint, oil, or grease. These coatings prevent the iron from oxidizing as long as the coating remains undamaged. Even small scratches can start the corrosion process. Also, an unreactive glass coating called enamel can be used in much the same way as paint or oil. Enamel is more resistant to damage.

GALVANIZING

A more advanced type of surface coating is applied during a process called galvanizing. In this process the iron is covered with a protective coating of zinc which prevents corrosion in two ways. First, it acts as a simple surface coating as discussed above, but unlike paint or enamel, if the surface becomes damaged, another method of protection exists. Zinc is more easily oxidized than iron, so if the potential for oxidation is there, the zinc will oxidize instead of the iron. Zinc acts as a sacrificial anode, allowing iron to be the cathode and not undergo oxidation.

CATHODIC PROTECTION

Similar to galvanizing, more reactive metals can simply be attached to the surface of the iron to act as sacrificial anodes. Magnesium, zinc, and aluminium are all used for this purpose. The sacrificial anode must be replaced once its ability to undergo oxidation is exhausted. This type of protection relies on other metals being more reactive to oxidation than iron. However, if iron is in contact with a metal that is less reactive, like tin or copper, the opposite effect is observed. In these cases, iron becomes the anode and because it has another metal as a redox partner, the oxidation occurs faster.

Practice

10. When comparing the rusting of iron to a galvanic cell, which component of the system acts as the salt bridge?

A. The iron

B. The iron (III) oxide

C. The rain water

D. There is no salt bridge

12.4.2.1 *use appropriate scientific vocabulary to communicate ideas related to electrochemistry*

When dealing with electrochemistry, a number of terms can easily be confused because many are related or even mean the same thing. It is important to understand how the terms interrelate, as it will allow for better communication and a fuller comprehension of the subject. The easiest way to accomplish this is to start out with a very broad term that encompasses all of the others and to break it down into more detailed parts.

Electrochemical reactions are often thought of as occurring in an electrochemical cell. Electrochemical cell is a term that refers to both galvanic and electrolytic cells. These two cells are distinguished by the galvanic cell representing a spontaneous reaction and the electrolytic cell representing a nonspontaneous one. The main parts of both types of cells are the cathode and anode. The reactions that occur at the cathode and anode are called half-reactions because they do not occur unless paired with another half-reaction.

All of electrochemisty is based on the principles of oxidation and reduction (or redox), with oxidation meaning the loss of electrons and reduction the gain of electrons. The half-reaction that occurs at the cathode is a reduction reaction and an oxidation reaction occurs at the anode. A chemical species that undergoes a reduction is called an oxidizing agent and one that undergoes an oxidation a reducing agent. This again goes back to the two types of reactions being linked.

In order to tell if an oxidation or reduction has occurred, it is necessary to examine the oxidation numbers of all atoms in the chemical equation. If an oxidation number increases, oxidation has occurred and vice versa for reduction. All of these terms are part of a hierarchy of the chemistry occurring in an electrochemical cell.

Practice

11. A chemical reaction is considered a redox reaction if

A. electrons are lost

B. electrons are gained

C. the reaction is spontaneous

D. electrons are lost by one atom and gained by another

12.4.2.2 *demonstrate oxidation-reduction reactions through experiments, and analyse these reactions*

EXPERIMENTS TO DETERMINE REDOX REACTIVITY

Oxidation and reduction can be experimentally observed by comparing the reactivity of one substance with another. This is based on the fact that certain metals, acids, and other compounds are better oxidizing or reducing agents than others. The easiest way to conduct these experiments is to use metal strips and observe what substances undergo reactions with them. In some cases, heating is necessary to speed the reaction up to an observable time scale.

Since the starting material is a solid metal, if a reaction occurs, it can be assumed that the metal is oxidized by the compound added to it and is therefore a better reducing agent than the compound. By repeated experiments, tables can be developed that organize different compounds according to their oxidizing and reducing ability. These redox tables are related to the activity series of metals and can be used to predict redox reactions. In addition, they follow the same trend as the table of standard reduction potentials, which can be used to determine if certain reactions will occur. Below are examples of three types of experiments that can be conducted.

REACTIVITY OF METALS WITH OTHER METALS

The easiest way to categorize the oxidizing ability of metals is to set up an experiment involving single replacement reactions. In these reactions, a solid metal is treated with a solutions of another metal salt. If a reaction occurs, it means that the ionic metal oxidized the solid metal. For instance, if solid magnesium is treated with aluminium chloride, the following reaction occurs:
$3Mg_{(s)} + 2AlCl_{3(aq)} \rightarrow 2Al_{(s)} + 3MgCl_{2(aq)}$.

This means that aluminium is a better oxidizing agent than magnesium. Repeating this for a variety of combinations of metals and solutions enables a table to be formed.

REACTIVITY OF OXYGEN WITH METALS

When certain metals are exposed to air, they react immediately and sometimes violently. This is usually the result of being oxidized by oxygen gas. Other metals react more slowly and some not at all. Experiments can be conducted where pure metals are exposed to oxygen and if oxygen is able to oxidize the metal, a metal oxide will form. In certain cases, the reaction is very slow and heating is required. In most cases, only the surface of the metal becomes oxidized. The more reactive a metal is with oxygen, the more likely it is to be oxidized by other metals.

REACTIVITY OF ACIDS WITH METALS

Solid metals can also undergo redox reactions with acids. With certain acids, the hydrogen ion released by the acid acts as the oxidizing agent by a reaction similar to the standard hydrogen electrode: $2H^{+}_{(aq)} + 2e^{-} \rightarrow H_{2(g)}$. However, there exist acids that are called oxidizing acids because this is not the reaction that occurs. In these acids, the anion product of the acid dissociation acids as the oxidizing agent. For instance, when copper reacts with nitric acid, the chemical equation is: $Cu_{(s)} + 4HNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2NO_{2(g)} + 2H_2O_{(l)}$. This means that nitrogen dioxide gas is evolved instead of hydrogen gas. The nitric acid underwent the reduction as it was a stronger oxidizing agent than H^{+} alone.

Identification of the gas evolved during the experiment can determine if whether or not the acid is an oxidizing acid. If no reaction occurs, it means that the neither dissociation product of the acid was a strong enough oxidizing agent.

Practice

12. Given the following table of experimental results, which reaction will not occur?

Metals	**$Cu_{(s)}$**	**$Fe_{(s)}$**	**$Pb_{(s)}$**
reacted with	none	$Cu^{2+}_{(aq)}$	$Cu^{2+}_{(aq)}$, $Fe^{3+}_{(aq)}$

A. $3Pb_{(s)} + 2Fe(NO_3)_{3(aq)}$
$\rightarrow 3Pb(NO_3)_{2(aq)} + 2Fe_{(s)}$

B. $Pb_{(s)} + Cu(NO_3)_{2(aq)}$
$\rightarrow Pb(NO_3)_{2(aq)} + Cu_{(s)}$

C. $3Cu_{(s)} + 2Fe(NO_3)_{3(aq)}$
$\rightarrow 3Cu(NO_3)_{2(aq)} + 2Fe_{(s)}$

D. $2Fe_{(s)} + 3Cu(NO_3)_{2(aq)}$
$\rightarrow 2Fe(NO_3)_{3(aq)} + 3Cu_{(s)}$

Numerical Response

Use the following information to answer the next question.

Three unknown metals are exposed to oxygen. When metal 1 is exposed, nothing happens. But upon heating, a thin oxide layer develops. Metal 2 immediately develops an oxide layer when exposed. Metal 3 never develops an oxide layer, even when heated.

13. Arrange the metals in order with the strongest reducing agent on the left.
____, ____, and ____

Open Response

Use the following information to answer the next question.

An experiment is conducted to determine the redox properties of copper, silver, zinc, and lead. The experimental setup includes multiple strips of each metal and solutions of $Cu(NO_3)_2$, $AgNO_3$, $Zn(NO_3)_2$, and $Pb(NO_3)_2$.

14. Write the redox half-reactions and the full single replacement reaction that occurs if silver nitrate reacts with zinc metal.

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12.4.2.3 *write balanced chemical equations for oxidation-reduction systems, including half-cell reactions*

BALANCING REDOX EQUATIONS

BALANCING HALF-REACTIONS

Balancing half-reactions is a relatively simple procedure. First the chemical species must be balanced and then the correct number of electrons must be added to balance the net charge.

For instance, in $Cl_2 + 2e^- \rightarrow 2Cl^-$, the net charge on each side is -2, and the chlorine is balanced as well. Special care must be taken when balancing half-reactions in acidic or basic solutions. However, simple steps can be followed for each situation. Take the example of balancing the half-reaction for the reduction of permanganate to manganese (II) ions.

Example

First, the unbalanced half-reaction must be written
$MnO_4^- \rightarrow Mn^{2+}$.

Next, balance the atoms that are not oxygen and hydrogen, which are already balanced.

Next, balance the oxygen atoms by adding water molecules (if in aqueous solution)
$MnO_4^- \rightarrow Mn^{2+} + 4H_2O$.

Since the reaction occurs in an acidic solutions, to balance the hydrogen atoms just add H^+
$MnO_4 + 8H^+ \rightarrow Mn^{2+} + 4H_2O$.

Finally, balance the charges by adding electrons
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$.

The steps are similar for balancing a half-reaction in a basic solution. The first three steps are the same so we can start with that point in the last example, after the hydrogens and oxygens are balanced as if in an acidic solution
$MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$.

To adjust for the basic conditions, add to both sides a number of hydroxide ions to equal the number of hydrogen ions in the equation
$MnO_4^- + 8H^+ + 8OH^- \rightarrow Mn^{2+} + 4H_2O + 8OH^-$.

Next, combine any hydrogen and hydroxide ions that are on the same side
$MnO_4^- + 8H_2O \rightarrow Mn^{2+} + 4H_2O + 8OH^-$.

Next, remove any water molecules present on both sides $MnO_4^- + 4H_2O \rightarrow Mn^{2+} + 8OH^-$.

It can be seen that the atoms are balanced, but the charge is not. The charge on the reactant side is −1 and on the products side is −6. Therefore, five electrons need to be added to the reactant side $MnO_4^- + 4H_2O + 5e^- \rightarrow Mn^{2+} + 8OH^-$.

FORMING BALANCED NET IONIC EQUATIONS FROM BALANCED HALF-REACTIONS

Full redox reactions can be balanced by taking the half-reactions and balancing them separately. To do this, it is important to identify which half-reaction is the oxidation half-reaction and which is the reduction. This can be done by looking at the oxidation numbers of atoms involved.

Then simply balance each half-reaction and determine the number of electrons involved in each. Multiply the half-reactions by the least common multiple of electrons to make them equal and add them. Then the electrons can be removed from both sides along with any atoms or molecules that are the same. Spectator ions can be added if the full chemical equation is required.

USING OXIDATION NUMBERS TO BALANCE EQUATIONS

Oxidation numbers can also be used to balance redox reactions. This is accomplished by insuring that the total increase of the oxidation number in the oxidized elements is equal to the total decrease in number of the reduced elements. Then complete the balancing by inspection. If in acid or basic conditions, the above steps can be followed to balance hydrogen and oxygen once the oxidation numbers are balanced.

Practice

15. Which values of x and y will balance the oxidation numbers in the following equation: $xH_2S + yO_2 \rightarrow SO_2 + H_2O$

A. 1 and 2　　**B.** 2 and 3

C. 3 and 4　　**D.** 6 and 12

Numerical Response

16. What is the coefficient of hydroxide in the balanced half-reaction of $IO_3^- \rightarrow I^-$ in a basic solution? ____

12.4.2.4 *determine oxidation and reduction half-cell reactions, direction of current flow, electrode polarity, cell potential, and ion movement in typical galvanic and electrolytic cells, including those assembled in the laboratory*

INTERPRETING ELECTROLYTIC CELLS

When presented with the components of a typical electrolytic or galvanic cell, it is important to be able to determine what half-reactions will drive the cell, and by doing so, make other conclusions about the cell. The first step is to identify anything in the cell setup that could undergo a redox reaction. This includes the electrodes and the aqueous solution (even water). It is also important to remember that acidic solutions contain H^+ and basic solutions contain OH^-, and certain half-reactions require the presence of either of these ions.

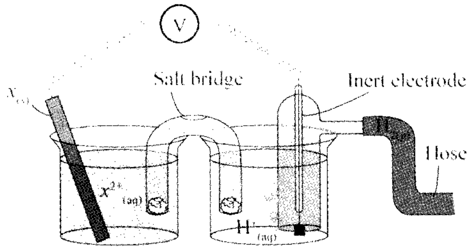
Once all the possibilities are identified, it is necessary to determine what is the strongest oxidizing agent present and what is the strongest reducing agent present by using the standard reduction potential table. The strongest oxidizing agent will have the highest standard reduction potential and will be located highest on the left (reduction) side of the table. The strongest reducing agent will have the lowest reduction potential and will appear lowest on the right (oxidation) side of the table.

The oxidation reaction will be read left to right and the reduction reaction will be read right to left. Everything else not involved in either of these half-reactions can be considered a spectator ion. Once the oxidation and reduction half-reactions have been chosen, the cell can be examined to see which side is the cathode and which is the anode. Remember that oxidation occurs at the anode and reduction occurs at the cathode. The electron flow always goes from anode to cathode.

In the case of a galvanic cell with a salt bridge, the non-reacting anions move to the anode and cations to the cathode to balance the charge. Cell potential is determined by the reduction potentials of the half-reactions. $E°_{\text{cell}} = E°_{\text{cathode}} - E°_{\text{anode}}$.

Practice

Use the following information to answer the next question.



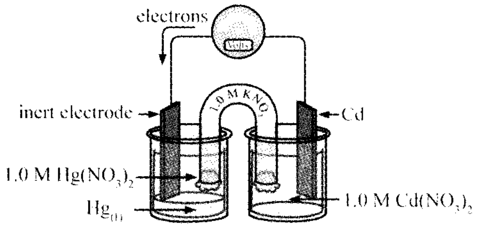
In this apparatus, the anions in the solution move from the hydrogen half-cell solution into the salt bridge and migrate toward the $X_{(s)}$ electrodes.

17. As this cell operates, electrons flow from

- **A.** $X_{(s)}$ to the inert electrode, and the pH in the hydrogen half-cell increases
- **B.** $X_{(s)}$ to the inert electrode, and the pH in the hydrogen half-cell decreases
- **C.** the inert electrode to $X_{(s)}$, and the pH in the hydrogen half-cell increases
- **D.** the inert electrode to $X_{(s)}$, and the pH in the hydrogen half-cell decreases

Numerical Response

Use the following information to answer the next question.

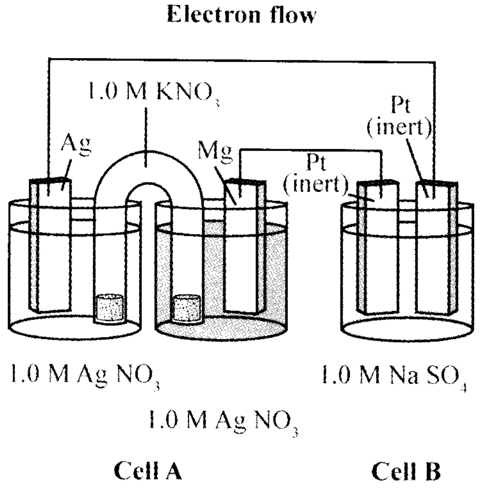


18. If the standard cell voltage is 1.25 V, what is the reduction half-cell potential for Cd? ____ V

Open Response

Use the following information to answer the next question.

Consider the following setup consisting of a galvanic cell connected to an electrolytic cell:

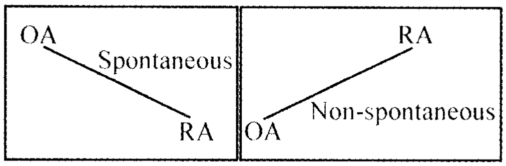


19. Explain how this cell will function. Please include a description of electron flow and the classification of the cathode and anode in each cell.

12.4.2.5 *predict the spontaneity of redox reactions and overall cell potentials by studying a table of half-cell reduction potentials*

PREDICTING REDOX REACTIONS

The spontaneity of redox reactions can be predicted by examining the relative positions of the oxidizing and reducing agents in a table of relative strengths. The table of standard reduction potentials is such a table. In this table, the oxidizing agents must be above the reducing agents for a spontaneous reaction. The figure below illustrates this:



In the standard reduction potential table, the strongest oxidizing agents are near the top, as these reactions have the highest potential to be reduced. This concept of spontaneity can be applied to electrochemical cells as well. In a galvanic cell, the cathode will be the half-reaction with the highest reduction potential. The anode can be thought of as the half-reaction with the highest oxidation potential. The oxidation potential can be thought of as the negative of the reduction potential. The formula that expresses this is $E_{\text{cell}}^{\circ} = E_{\text{red}}^{\circ} + E_{\text{ox}}^{\circ}$.

Practice

20. Which of the following describes a spontaneous electrochemical reaction?

A. A negative cell potential

B. The strongest oxidizing agent has a lower reduction potential than the strongest reducing agent

C. The cathode half-reaction has a reduction potential of 0.85 V and the anode has a reduction potential of 1.1 V

D. The oxidation potential of the anode is 0.36 V and the reduction potential of the cathode is −0.13 V

Open Response

21. Explain whether a copper or zinc pipe would be the **best** for transporting hydrochloric acid.

12.4.2.7 *measure through experimentation the mass of metal deposited by electroplating, and apply Faraday's law to relate the mass of metal deposited to the amount of charge passed*

Electroplating is a process where an object is coated with a thin layer of metal using an electrolytic cell. The object to be coated is the cathode and the thin layer of metal comes from the reduction of metal ions in solution to solid metal. The anode is usually the solid state of the desired metal coating and acts as a source of metal ions by undergoing oxidation. An external power source is required as it is an electrolytic process. The electrolyte solution is usually composed of metal salts of the desired metal coating. The cathode itself does not undergo any redox chemistry and acts as an inert electrode. The metal plating often acts to protect the object or to improve the appearance of the object. Plating is often useful for circuit components to improve electrical properties. Common types of plating include nickel, copper, chrome, zinc, gold, silver, and platinum. Plating follows Faraday's law in the same way as other electrolytic processes.

Practice

22. A silver ornament is plated with gold. The electrolyte used in this process is

A. $NiSO_4$ solution

B. K_2SO_4 solution

C. $K[Au(CN)_2]$ solution

D. $K[Ag(CN)_2]$ solution

Numerical Response

Use the following information to answer the next question.

A piece of jewellery needs to be electroplated with gold. However, the cost must not exceed twenty dollars and the price of gold is 32.53 dollars per gram.

23. How many seconds must a 20.0 A current be applied to equal twenty dollars of gold? ____ s (Give answer to **three** significant figures)

Open Response

Use the following information to answer the next question.

A penny weighs 2.52 g before an electroplating experiment. During the experiment, it is electroplated with zinc and weighs 2.74 g after. The electroplating procedure took only 20.0 s.

24. What current was applied during the electroplating? Show all work.

12.4.3.1 *describe examples of common galvanic cells and evaluate their environmental and social impact*

Commercial Galvanic Cells

The most prevalent use of galvanic cells in the commercial world is for batteries. A battery is simply a set of galvanic cells combined in series. The individual cell potentials combine to give the battery its full voltage. The battery is designed so that no electrochemical reactions occur until the circuit is complete, meaning that the battery is connected to the device it needs to power. Below are some examples of different battery designs.

Primary Batteries

Disposable batteries are mostly dry cell batteries. A dry cell is a galvanic cell where the electrolyte is a paste instead of a liquid. Regardless of the design of the cell, the galvanic cell principles are still at play, with the separated half-reactions providing the cell potential that allows the battery to generate a current. Disposable batteries cease to provide power when the reactants are used up.

Dry cells contain a zinc anode and an inert carbon cathode. The electrolytic paste is made up of manganese (IV) oxide, zinc chloride, ammonium chloride, and soot carbon. The reaction that occurs at the zinc anode is $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$ while at the inert cathode, the reaction is
$2MnO_{2(s)} + H_2O_{(l)} + 2e^- \rightarrow Mn_2O_{3(s)} + 2OH^-_{(aq)}$.

A longer lasting version of the dry cell battery is the alkaline cell battery. In these batteries, the ammonium chloride and zinc chloride are replaced by potassium hydroxide. This changes the anode reaction to
$Zn_{(s)} + 2OH^-_{(aq)} \rightarrow ZnO_{(s)} + H_2O_{(l)} + 2e$ and the cathode reaction to
$MnO_{2(s)} + 2H_2O_{(l)} + 2e^- \rightarrow Mn(OH)_{2(s)} + 2OH^-_{(aq)}$.

Finally, smaller batteries designed for watches and pacemakers are called button cell batteries. One example, called a mercury button battery, has a zinc container that acts as an anode and an inert stainless steel cathode. The alkaline electrolytic paste has mercury oxide. The anode reaction is the same as the standard alkaline battery, but the reaction at the cathode is
$HgO_{(s)} + H_2O_{(l)} + 2e^- \rightarrow Hg_{(l)} + 2OH^-_{(aq)}$.

Batteries have had a major impact on society by allowing for portable electronics. The button cell battery has been extremely important in this because its small design makes way for even smaller electronics. Human health has also been greatly impacted with the development of pacemakers and hearing aids. However, there are some concerns about the batteries being disposable and releasing polluting chemicals into the environment.

Secondary Batteries

Secondary batteries are galvanic cells that can be recharged when the reactants are used up. This is done by temporarily converting the galvanic cell into an electrolytic cell and thereby reversing the electrochemical reaction. This means that reactants become products and products become reactants. Two common types of secondary batteries are nickel-cadmium and lead-acid. Lead-acid batteries are used in cars and are recharged by the alternator.

The discharging reaction is
$Pb_{(s)} + PbO_{2(s)} + 2H_2SO_{4(aq)} \rightarrow 2PbSO_{4(s)} + 2H_2O_{(l)}$
while the recharging reaction is the exact reverse.

Secondary batteries provide a way to limit the number of batteries that end up in landfills. However, side reactions and the breakdown of products and reactants over time means that eventually rechargeable batteries must be replaced as well.

FUEL CELLS

Fuel cells are unlike batteries in that they have a continuous supply of reactants to keep the cell producing electricity. This means that the cell can operate indefinitely. Fuel cells are important because they don't involve the burning of fossil fuel and the production of green house gases. In addition, they are more efficient than traditional methods of producing electricity.

The current design of the hydrogen fuel cell uses the reaction between oxygen and hydrogen gas to produce water. The cathode reaction is $O_{2(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^-_{(aq)}$ and the anode reaction is $2H_{2(g)} + 4OH^-_{(aq)} \rightarrow 4H_2O_{(l)} + 4e^-$.

Practice

Open Response

25. What alternator potential is required to recharge a lead-acid battery?

12.4.3.2 *explain how electrolytic processes are involved in industrial processes*

Electrolysis is useful for industry because it forces nonspontaneous reactions to occur. Most products of spontaneous reactions can be collected directly from the environment, as a spontaneous reaction is likely to occur naturally. However, some very useful products are not available through spontaneous reactions. The most used application of electrolysis in industry are for the generation of useful elements and metal refining.

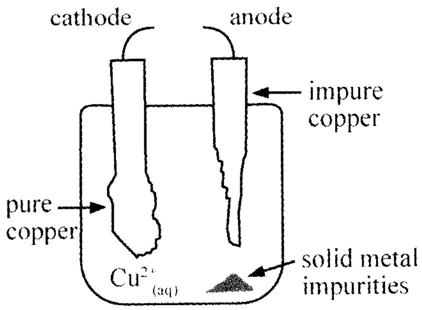
ELECTROLYSIS FOR ELEMENT PRODUCTION

Metals that have very negative reduction potentials are usually do not occur in nature because of their ease of oxidation. This means that in order to obtain them, electrolysis is necessary. Examples of such metals are those at the bottom of the standard reduction potential table. Sodium, zinc, magnesium and strontium are a few examples.

In addition, certain oxidation reactions are desirable that don't occur naturally for the same reason. The production of chlorine gas from chloride ions has a very low oxidation potential which means electrolysis is required. The production of aluminium from aluminium oxide (bauxite) is another important use of electrolysis.

ELECTROLYSIS FOR METAL REFINING

Electrolysis can also be used to refine a metal that is too impure for industrial processes. Electrorefining uses electrolysis to deposit a pure metal at the cathode and an impure metal at the anode. This technique is commonly used on copper because impurities decrease its conductivity. Copper is widely used in electronics as a conductor. The process works by using a very pure sample of copper as the cathode and impure copper as the anode. Therefore, both the copper and impurities are oxidized at the anode but only copper is reduced at the cathode. The figure below is a simple representation of the process:



Practice

Open Response

26. How many moles of electrons are required to reduce 1 mole of bauxite?

12.4.3.3 *research and assess environmental, health, and safety issues involving electrochemistry*

Electrochemical processes have many impacts on our environment and health. Corrosion of iron structures is responsible for billions of dollars in repair cost each year. In addition, the weakening of these structures by corrosion can pose serious risks to human health. Iron is used in buildings, automobiles, aircraft, and ships, all of which can cause injury or death if structurally compromised.

While corrosion is an example of how electrochemistry can negatively affect human health and safety, many other applications of electrochemistry are helpful. For instance, the chlorine produced from the Downs cell and the chlor-alkali process can be used to purify water for human consumption. In the chlor-alkali process, brine is electrolyzed to produce chlorine gas, hydrogen gas, and sodium hydroxide. The hydrogen gas can be used in hydrogen fuel cells, another application of electrochemistry with possible vast implications on our environment.

Hydrogen fuel cells have the potential to replace combustion engines in cars, thereby reducing the amount of carbon dioxide released into the atmosphere. The chlorine can be mixed with water and sodium hypochlorite to produces the antibacterial agent hypochlorous acid. However, there are issues concerning the use of chlorinated water, because hypochlorous acid can react with certain organic compounds to produce toxic substances.

Practice

27. Which of the following is true about the chlor-alkali process?
    - **A.** Chloride ions are reduced to form chlorine gas
    - **B.** Solid sodium is produced at the cathode
    - **C.** Water is reduced at the cathode
    - **D.** The reaction is spontaneous

SOLUTIONS–ELECTROCHEMISTRY

1. A	7. A	13. 2,1,3	19. OR	25. OR
2. C	8. OR	14. OR	20. D	26. OR
3. C	9. A	15. B	21. OR	27. C
4. 3	10. D	16. 6	22. C	
5. D	11. D	17. A	23. 45.2	
6. OR	12. C	18. –0.40	24. OR	

1. A

Carbon will only have a negative oxidation number if it is bonded to atoms that normally have a positive oxidation number. In methane, the oxidation number of carbon is –4 because each hydrogen has an oxidation number of +1.

2. C

While the designs of galvanic and electrolytic cells are different, they both have cathodes, anodes, and oxidizing and reduction agents. The major difference is the spontaneity of the redox reaction. This is what causes the galvanic cell to produce power and the electrolytic cell to require power.

3. C

In order to have a galvanic cell that functions, meaning it has a positive cell potential, half-cell reduction must occur at one and oxidation at the other. In addition, the half-cell potential must be different from one another, which means they must involve different compounds. Because the given half-cell is the oxidation of element A, the second half cell must be the reduction of something else. This eliminates choices A and B. C is a reduction of species B, which therefore satisfies all the necessary conditions.

4. 3

$AuCl_3$ in an aqueous solution splits into the Au^{3+} ion and Cl^- ion, while AgCl splits into Ag^+ and Cl^-. Therefore, in one half-cell, the half-reaction is a three electron reduction of the gold ion into solid gold and in the other half-cell, the half-reaction is the one electron oxidation of silver to Ag^+.

The overall balanced net redox reaction is as follows:

$$Au^{3+}_{(aq)} + 3e^- \rightarrow Au_{(s)}$$

$$\underline{Ag_{(s)} \rightarrow Ag^+_{(aq)} + 3e^-}$$

$$Ag_{(s)} + Au^{3+}_{(aq)} \rightarrow Au_{(s)} + Ag^+_{(aq)}$$

This means the overall redox reaction involves 3 electrons.

5. D

Because the cell potential is a negative value, it means that the voltmeter is hooked up backwards and that hydrogen should be the cathode and chromium the anode. Thus hydrogen undergoes the reduction and chromium undergoes the oxidation.

6. Open Response

The standard iron-lead cell involves the half-reaction $Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$ with standard reduction potential of –0.45 V and $Pb^{2+}_{(aq)} + 2e^- \rightarrow Pb_{(s)}$ with a standard reduction potential of –0.13 V. Because the lead half-reaction is higher on the standard reduction potential chart, it will be the cathode and the iron electrode will be the anode. The cell notation is $Fe_{(s)}|\ Fe^{2+}_{(aq)}||\ Pb^{2+}_{(aq)}|Pb_{(s)}$ and the standard cell potential is $\Delta E° = -0.13\ V - (-0.45\ V) = 0.32\ V$.

7. A

One Faraday is the charge of a mole of electrons: $\frac{1.602 \times 10^{-19}\ C}{1e^-} \times \frac{6.022 \times 10^{23}\ e^-}{1\ mol}$

$= 9.647 \times 10^4\ C/mol$

8. Open Response

To relate moles of gas to volume, use the ideal gas law: $V = \frac{nRT}{P}$. Then all that is left to do is substitute in an expression for the number of moles based on current and time: $n = \frac{\text{current} \times \text{time}}{9.647 \times 10^4\ C/mol}$. Now combine the two: $V = \frac{(\text{current} \times \text{time})RT}{P(9.647 \times 10^4\ C/mol)}$.

9. A

The silver nitrate solution would undergo the following redox half-reaction: $Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$ which means one electron produces one atom of silver. All of the other nitrates involve more than one electron to produce an atom of solid product. Therefore, the same current being applied for the same amount of time to each nitrate solution would result in the highest yield of silver.

10. D

In rusting, the salt bridge does not exist. This allows the products of the oxidation and reduction reactions to further combine into iron (II) hydroxide. After this, the iron (II) hydroxide can undergo additional oxidation to make iron (III) hydroxide, which breaks down into the hydrated iron (III) oxide (the rust).

11. D

Redox implies the oxidation of one atom and the reduction of another. Both reactions are required to give the electrons the potential to transfer.

12. C

The table indicates that copper is the least reactive and lead is the most. Therefore, lead has the greatest tendency to lose electrons and is therefore the strongest reducing agent. Conversely, Cu^{2+} is the strongest oxidizing agent and Pb^{2+} is the weakest. In order to replace a metal in a single replacement reaction, the solid metal must be a stronger reducing agent then the solid of the metal to be replaced. Therefore, copper cannot replace Pb^{2+} or Fe^{3+}.

13. 2,1,3

The strongest reducing agent is going to be the most easily oxidized. The metal that reacts with oxygen immediately is the best reducing agent and the one that does not react at all is the worst.

14. Open Response

$Ag^{+}_{(aq)} + e^{-} \rightarrow Ag_{(s)}$ and $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^{-}$ are the half reactions. The full single displacement reaction is:

$Zn_{(s)} + 2AgNO_{3(aq)} \rightarrow Zn(NO_3)_{2(aq)} + 2Ag_{(s)}$

15. B

The sulfur atom is oxidized from a −2to a +4 state. Oxygen is reduced from zero to −2. Since there are two oxygen atom from each O_2 molecule, this represents a total change of 4 electrons, whereas the sulfur represents a change of 6 electrons. 12 is the least common multiple of 6 and 4 so x is 2 and y is 3.

16. 6

First, notice that everything but oxygen is already balanced. Then, balance oxygen by adding water

$IO_3^{-} \rightarrow I^{-} + 3H_2O$.

Next, balance hydrogen by adding H^{+},

$IO_3^{-} + 6H^{+} \rightarrow I^{-} + 3H_2O$. Since the solution is basic, hydroxide can be added to both sides to cancel the hydrogen ion $IO_3^{-} + 6H_2O \rightarrow I^{-} + 3H_2O + 6OH^{-}$.

Cancel the waters and balance the charge

$IO_3^{-} + 3H_2O + 6e^{-} \rightarrow I^{-} + 6OH^{-}$.

17. A

In all electrochemical cells, the anions flow through the electrolyte to the anode. Evidently, the anode in the cell shown is the $X_{(s)}/X^{2+}_{(aq)}$ half-cell.

Thus, the hydrogen half-cell is the cathode, and the half-reaction occurring there will be the reduction of $H^{+}_{(aq)}$, as follows:

$$2H^{+}_{(aq)} + 2e^{-} \rightarrow H_{2(g)} \quad E^{\circ}_{RC} = 0.00V$$

With fewer $H^{+}_{(aq)}$ ions in the cathode as the cell operates, the cathode pH is destined to rise (a less acidic half-cell). In all cells, since electrons flow from the anode to the cathode, it is obvious that in the cell shown, electrons will flow from $X_{(s)}$ to the inert electrode, and pH in the hydrogen half-cell will increase.

18. −0.40

For all electrochemical cells, the standard cell electrochemical potential is the difference between the cathode and the anode electrode potentials. As follows:

$E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode}$

For this cell,

$E^{\circ}_{cathode} = +0.80\ V(Hg \mid Hg^{2+})$ and $E^{\circ}_{cell} = 1.25\ V$

Therefore, $E^{\circ}_{anode} = E^{\circ}_{cathode} - E^{\circ}_{cell}$

$= +0.80\ V - 1.25\ V = -0.40\ V$.

19. Open Response

In this cell, the electrochemical potential in cell A is used to drive the electrolytic reaction in cell B. In cell A, silver is the cathode and magnesium is the anode because the silver reduction half-reaction has a higher reduction potential than the magnesium half-reaction. Therefore, electrons will flow from the magnesium electrode and enter cell B. Because of the presence of electrons at the inert electrode connected to the anode of cell A, reduction will occur there making it the cathode. In order to continue the flow of electrons back to the cathode of cell A, oxidation must occur at the inert electrode connected to the silver electrode. this would make it the anode of cell B.

20. D

The cell potential can be thought of as the oxidation potential of the anode added to the reduction potential of the cathode. This means that choice D gives a positive cell potential, and therefore a spontaneous reaction.

21. Open Response

To transport a material, the pipe should not spontaneously react with what it is transporting. Since the pipe would be either solid zinc or copper, oxidation would occur there. By the spontaneity rule, the reducing agent (the pipe) should be above the oxidizing agent H^{+} in the standard reduction potential table. Copper is above and zinc is below, so copper would be the better choice.

22. C

The metal to be electroplated is the cathode. The metal to be deposited is the anode. A soluble metal salt is used as an electrolyte. For gold plating, pure gold plate is used as the anode, silver ornament as the cathode, soluble gold salt, potassium dicyanoaurate $K[Au(CN)_2]$, as the electrolyte.

23. 45.2

First, find out how many grams twenty dollars is equivalent to $20 \text{ dollars} \times \frac{1 \text{ g}}{32.53 \text{ dollars}} = 0.615 \text{ g}$.

Now find the number of moles of gold to which that is equal: $0.615 \text{ g} \times \frac{1 \text{ mol}}{196.97 \text{ g}} = 3.12 \times 10^{-3} \text{ mol}$.

The half-reaction for the reduction of gold is $Au^{3+}_{(aq)} + 3e^- \rightarrow Au_{(s)}$. So three moles of electrons are needed for every mole of gold:

$3.12 \times 10^{-3} \text{ mol Au} \times \frac{3 \text{ mol e}^-}{1 \text{ mol Au}} = 9.36 \text{ mol e}^-$.

This translates into a charge of

$9.36 \times 10^{-3} \text{ mol} \times 9.65 \times 10^4 \frac{\text{C}}{\text{mol}} = 9.03 \times 10^2 \text{ C}$. To convert to time, $9.03 \times 10^2 \text{ C} \times \frac{1 \text{ s}}{20.0 \text{ C}} = 45.2 \text{ s}$.

24. Open Response

The difference in the mass of the penny before electroplating and after is the weight of zinc added.
$2.74 \text{ g} - 2.52 \text{ g} = 0.22 \text{ g}$, which is
$0.22 \text{ g} \times \frac{1 \text{ mol}}{65.39 \text{ g}} = 3.36 \times 10^{-3} \text{ mol}$. The half-reaction for zinc reduction is $Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)}$ so for every mole of zinc, two moles of electrons are required.
To calculate the charge needed
$6.72 \times 10^{-3} \text{ mol} \times 9.647 \times 10^4 \text{ C}/(\text{mol})$
$= 6.49 \times 10^2 \text{ C}$, which applied over twenty seconds is
$\frac{6.49 \times 10^2 \text{ C}}{20.0 \text{ s}} = 32.5 \text{ A}$.

25. Open Response

The cathode reduction potential is 1.69 V and the anode reduction potential is −0.36 V.
This means that the cell potential is 2.05 V. Therefore, to do the reverse reaction would require that much voltage.

26. Open Response

Bauxite has the formula $Al_2O_3 \cdot xH_2O$. The oxidation state of aluminium is +3. So in order for every mole of bauxite to be fully reduced, six moles of electrons are required.

27. C

The chlor-alkali process involves the oxidation of chloride and the reduction of water to produce chlorine and hydrogen gas. It is an electrolytic process, so the reaction is not spontaneous.

Self Test

1. In the following redox chemical equation, which species is being reduced?

$2Ag^{+}_{(aq)} + Cu_{(s)} \rightarrow 2Ag_{(s)} + Cu^{2+}_{(aq)}$

A. Ag^{+} **B.** Ag

C. Cu^{2+} **D.** Cu

2. In a galvanic cell, what best describes the function of the salt bridge?

A. To refill the electrolyte solution

B. To provide a path for redox reactions

C. To provide ions for reactions

D. To balance charge

Numerical Response

Use the following information to answer the next question.

Three half-cells are constructed. The first contains a cobalt electrode in a $Co(NO_3)_3$ solution, the second has a silver electrode in a $AgNO_3$ solution, and the third has a copper electrode in a $CuNO_3$ solution. The half-cells can be combined with one another and the cell-potential can be read on a voltmeter. When the cobalt and copper half-cells are combined, the voltmeter reads 1.28 V. When the silver and the copper half-cells are combined, the voltmeter reads 0.33 V. It is known that the half reaction for copper has the lowest reduction potential.

3. What will the voltmeter read when the cobalt and silver half-cells are combined? ____ V

4. What would the standard reduction potential be of the hydrogen electrode if all standard values were measured against a standard nickel electrode?

A. 0.01 V **B.** −0.26 V

C. 0.26 V **D.** −0.01 V

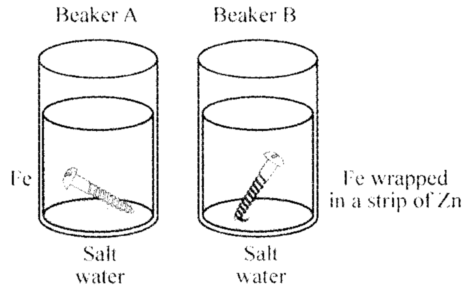
Open Response

5. How many hours does a current of 500.0 mA need to be applied to a solution of silver nitrate to produce 0.10 mol of solid silver? Show all work.

Open Response

Use the following information to answer the next question.

In the following experimental setup, one iron nail is covered in zinc while the other is not.



6. Explain what will happen to the two nails over time in terms of redox reactions and corrosion.

Use the following information to answer the next question.

An experiment is conducted to determine the redox properties of copper, silver, zinc, and lead. The experimental setup includes multiple strips of each metal and solutions of $Cu(NO_3)_2$, $AgNO_3$, $Zn(NO_3)_2$, and $Pb(NO_3)_2$.

7. When a drop of copper nitrate is placed on the lead strip, a reaction is observed but not when it is placed on the silver strip. Which of the following statements is true?

A. Pb^{2+} is a better oxidizing agent than Cu^{2+}

B. Cu^{2+} is a better oxidizing agent than Ag^{+}

C. Pb^{2+} is a better oxidizing agent than Ag^{+}

D. Cu^{2+} is a better oxidizing agent than Pb^{2+}

Open Response

8. Make a standard reduction potential table with three half-reactions that summarizes the results of an experiment where aluminium reacted with hydrochloric acid, but copper did not.

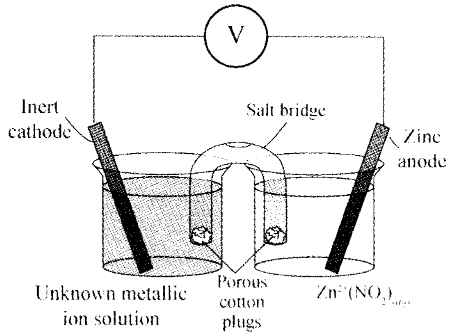
Open Response

9. Balance the following redox equation that takes place in an acidic solution:

$Cr_2O_7^{2-} + Fe \rightarrow Cr^{3+} + Fe^{2+}$

Use the following information to answer the next question.

To determine the identity of an unknown metallic ion in a solution, a student designed the voltaic cell shown below.



10. If the cell generates a voltage of +1.24 V under standard conditions, the half-reaction occurring at the cathode will have an electrode potential of

A. +2.00 V	**B.** −2.00 V
C. +0.48 V	**D.** −0.48 V

11. Under standard conditions, what voltage is needed to initiate the electrolysis of molten sodium chloride to solid sodium and chlorine gas with inert electrodes?

A. −4.07	**B.** 4.07
C. −1.35	**D.** 1.35

Open Response

12. In the electrolysis of molten sodium chloride, how many grams of solid sodium and chlorine gas are produced if a current of 1.00 A is applied for 20.0 min? Show all work.

Numerical Response

13. What mass of chromium can be plated onto a car bumper if a current of 2 000.0 A is applied for 10.00 min? ____ g

SOLUTIONS

1. A	**5. OR**	**9. OR**	**13. 215.5**
2. D	**6. OR**	**10. C**	
3. 0.95	**7. D**	**11. B**	
4. C	**8. OR**	**12. OR**	

1. A

The charge of the silver ion goes from 1+ to 0 which represents the addition of one electron for each silver ion. The addition of electrons is a reduction.

2. D

The salt bridge acts to counterbalance the difference in charge that builds up at the cathode and anode due to the transfer of electrons. Unreactive positive ions in the salt bridge move to the cathode and negative ions move to the anode. Without the salt bridge, the building of negative charge at the cathode would eventually prevent more electrons from travelling there.

3. 0.95

Since the cell potential is the difference in the reduction potentials of the half-reactions and it can be assumed that copper is the anode, Co – Cu = 1.28and Ag – Cu = 0.33. If we want to find the combination of Co and Ag, simple substitution can be used: Co – (Ag – 0.33) = 1.28 so Co – Ag = 0.95. Because the answer came out positive, silver is indeed the anode. If it had been negative, it would have just meant that cobalt was the anode.

4. C

Nickel's reduction potential is –0.26 V when compared with the standard hydrogen electrode. If all the reduction potentials were measured against nickel, they would be 0.26 V higher than on the current table. This would mean that the standard reduction potential for hydrogen would be 0.26 V.

5. Open Response

The redox half reaction is $Ag^{+}_{(aq)} + e^{-} \rightarrow Ag_{(s)}$ so 0.10 mol of electrons are need to produce 0.10 mol of silver. To get the necessary moles of electrons

$9.647 \times 10^{4}\ C/mol \times 0.10\ mol = 9.6 \times 10^{3}\ C.$

At 0.5000 C / s the current needs to be running for

$9.6 \times 10^{3}\ C \times \frac{s}{0.5000\ C} = 1.9 \times 10^{4}\ s$ which is

$1.9 \times 10^{4}\ s \times \frac{h}{3\,600\ s} = 5.4\ h.$

6. Open Response

The iron nails are both in a situation where rusting can occur because oxygen, water, and an electrolyte are all present. However, the second iron nail is covered in zinc, much in the same way as galvanized iron. Therefore, zinc, being the more easily oxidized metal, will act as the anode and iron as the cathode which will prevent rusting until all the zinc is oxidized. The first nail has no such protection and will begin to rust immediately.

7. D

By reacting with the lead strip, copper is more reactive than lead. However, since there was no reaction with silver, it means that copper is not as reactive as silver. This means that Cu^{2+} can oxidize $Pb_{(s)}$ but not $Ag_{(s)}$. From this information, it can be inferred that Ag^{+} is the best oxidizing agent of the three and Pb^{2+} is the worst.

8. Open Response

Since the reaction of aluminium and hydrochloric acid implies that H^{+} was able to oxidize aluminium to form hydrogen gas, it means that the H^{+} reduction half-reaction has a higher reduction potential than the reduction half-reactions of the aluminium. However, by not reacting with copper, it means that the copper reduction half-reaction has a higher reduction potential than the H^{+}. The table is below:

$Cu^{2+}_{(aq)} + 2e^{-} \rightleftharpoons Cu_{(s)}$
$2H^{+}_{(aq)} + 2e^{-} \rightleftharpoons H_{2(g)}$
$Al^{3+}_{(aq)} + 3e^{-} \rightleftharpoons Al_{(s)}$

9. Open Response

The unbalanced half-reactions are $Cr_2O_7^{2-} \rightarrow Cr^{3+}$ and $Fe \rightarrow Fe^{2+}$.

The oxidation half-reaction is balanced simply by adding electrons $Fe \rightarrow Fe^{2+} + 2e^{-}$.

The reduction requires more steps. First, balance the chromium $Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$ then the oxygen by adding water $Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$ and then hydrogen $Cr_2O_7^{2-} + 14H^{+} \rightarrow 2Cr^{3+} + 7H_2O$ and finally charge $Cr_2O_7^{2-} + 14H^{+} + 6e^{-} \rightarrow 2Cr^{3+} + 7H_2O$.

Six is the lowest common multiple, so the final equation is three times the oxidation equation added to the balanced reduction equation

$Cr_2O_7^{2-} + 14H^+ + 3Fe \rightarrow 2Cr^{3+} + 3Fe^{2+} + 7H_2O$.

10. C

$E^\circ_{net} = E^\circ_{r(cathode)} - E^\circ_{r(anode)}$

$+1.24\ V = E^\circ_{r(cathode)} - (-0.76\ V)$

$+0.48\ V = E^\circ_{r(cathode)}$

11. B

The two half-reactions involved are $2Cl^- \rightarrow Cl_2 + 2e^-$ with a reduction potential of 1.36 V and $Na^+ + e^- \rightarrow Na$ with a reduction potential of −2.71 V. Since reduction takes place at the cathode, the necessary voltage will be $-2.71 - (1.36) = -4.07$. This means a voltage of 4.07 V will need to be applied.

12. Open Response

The reaction at the cathode is $Na^+_{(l)} + e^- \rightarrow Na_{(s)}$ and at the anode $2Cl^-_{(l)} \rightarrow Cl_{2(g)} + 2e^-$. The first step is to find the amount of charge based on the current and time:

$1.00 C/s \times 20.0\ \text{min} \times \frac{60\ \text{s}}{1\ \text{min}} = 1.20 \times 10^3\ C$. Next, find how many moles of electrons that corresponds to:

$1.20 \times 10^3\ C \times \frac{1\ \text{mol}}{9.647 \times 10^4\ C} = 1.24 \times 10^{-2}\ \text{mol}$.

The relationship between solid sodium and moles of electrons is 1 to 1 and for chlorine gas to electrons is 1 to 2. Therefore, the mass of solid sodium is

$1.24 \times 10^{-2}\ \text{mol} \times \frac{22.99\ \text{g}}{\text{mol}} = 0.286\ \text{g}$ and for chlorine gas $6.22 \times 10^{-3}\ \text{mol} \times \frac{70.9\ \text{g}}{\text{mol}} = 0.441\ \text{g}$.

13. 215.5

The half-reaction at the cathode is $Cr^{3+}_{(aq)} + 3e^- \rightarrow Cr_{(s)}$. The charge generated by the current and time is

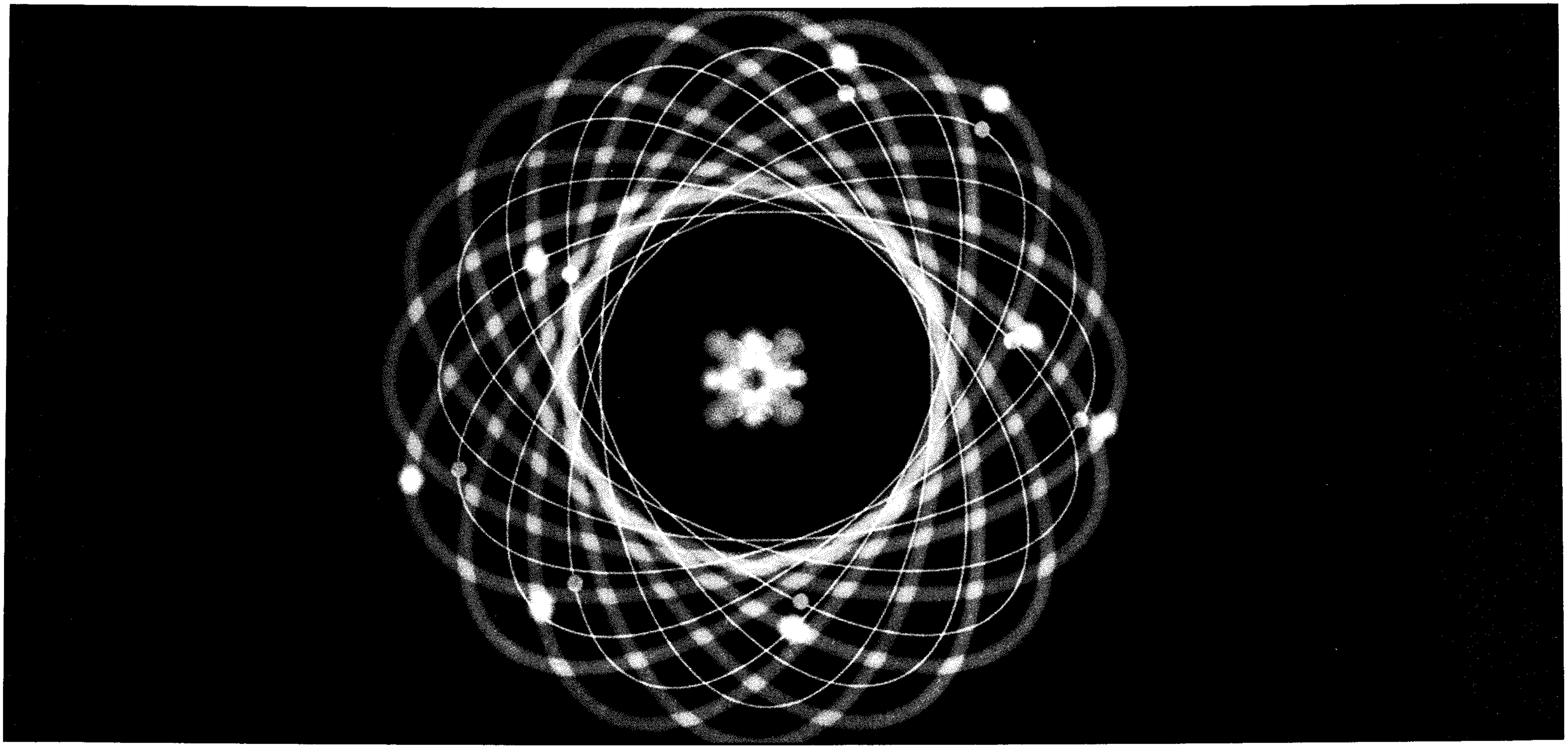
$2\,000.0\ C/s \times 10.00\ \text{min} \times \frac{60\ \text{s}}{1\ \text{min}} = 1.200 \times 10^6\ C$

which translates to $1.20 \times 10^6\ C \times \frac{1\ \text{mol}}{9.65 \times 10^4\ C}$

$= 12.44$ mol of electrons. Since three electrons are required for every mole of chromium $\frac{12.44\ \text{mol}}{3}$

$= 4.15\ \text{mol} \times \frac{52.00\ \text{g}}{\text{mol}} = 215.5\ \text{g}$.

Structure and Properties



Structure and Properties

Table of Correlations

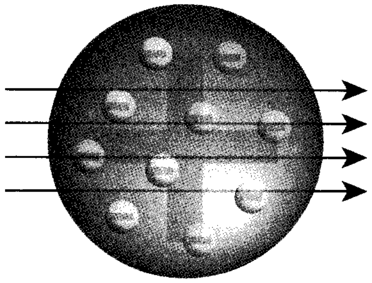
Specific Expectation	Practice Questions	Self Test Questions
12.5.1 Understanding Basic Concepts		
12.5.1.1 *explain the experimental observations and inferences made by Rutherford and Bohr in developing the planetary model of the hydrogen atom*	1	1
12.5.1.2 *describe the quantum mechanical model of the atom and the contributions of individuals to this model*	2, 3	2
12.5.1.3 *list characteristics of the s, p, d, and f blocks of elements, and explain the relationship between position of elements in the periodic table, their properties, and their electron configurations*	4a, 4b, 4c, 4d, 4e, 4f, 5	3
12.5.1.4 *explain how the properties of a solid or liquid depend on the nature of the particles present and the types of forces between them*	6, 7a, 7b, 7c, 7d, 7e, 7f	4
12.5.1.5 *explain how the Valence Shell Electron Pair Repulsion (VSEPR) model can be used to predict molecular shape*	8, 9	5
12.5.2 Developing Skills of Inquiry and Communication		
12.5.2.1 *use appropriate scientific vocabulary to communicate ideas related to structure and bonding*	10	
12.5.2.2 *write electron configurations for elements in the periodic table, using the Pauli exclusion principle and Hund's rule*	11, 12	6, 7
12.5.2.3 *predict molecular shape for simple molecules and ions, using the VSEPR model*	13, 14a, 14b	8a, 8b
12.5.2.4 *predict the polarity of various substances, using molecular shape and the electronegativity values of the elements of the substances*	15, 16, 17	9, 10
12.5.2.5 *predict the type of solid (ionic, molecular, covalent network, or metallic) formed by a substance, and describe its properties*	18	11
12.5.2.6 *conduct experiments to observe and analyse the physical properties of different substances, and to determine the type of bonding present*	19, 20	12
12.5.3 Relating Science to Technology, Society, and the Environment		
12.5.3.1 *describe some applications of principles relating to atomic and molecular structure in analytical chemistry and medical diagnosis*	21	
12.5.3.2 *describe some specialized new materials that have been created on the basis of the findings of research on the structure of matter, chemical bonding, and other properties of matter*	22a, 22b	
12.5.3.3 *describe advances in Canadian research on atomic and molecular theory*	23	

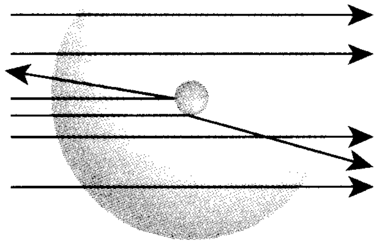
12.5.1.1 *explain the experimental observations and inferences made by Rutherford and Bohr in developing the planetary model of the hydrogen atom*

RUTHERFORD'S ATOMIC MODEL

In Rutherford's gold foil scattering experiment, he found that alpha particles scatter much more than a solid model of the atom would predict. He concluded that almost all of the mass and all the positive charge of an element are located in the centre of an atom or nucleus, and the negative charge is located with the electrons orbiting the nucleus like planets orbit the sun.

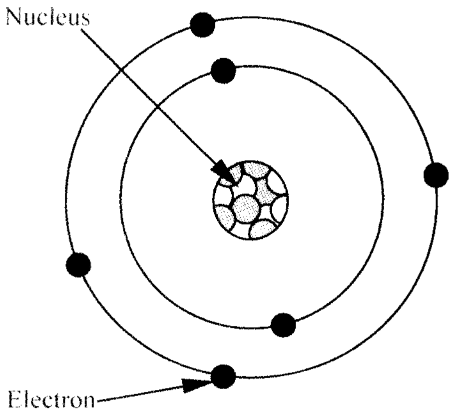
The following two diagrams show first what would happen to an atom that behaved in accordance with a solid atomic model, and then how atoms of gold behaved in Rutherford's experiment.





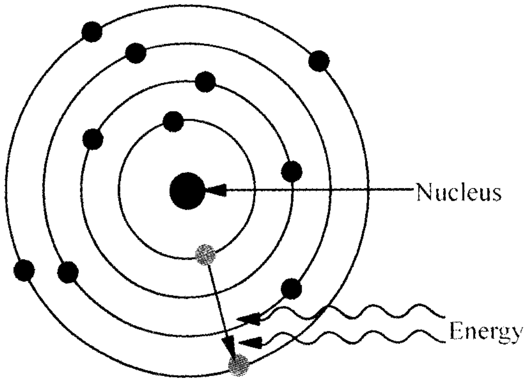
BOHR'S ATOMIC MODEL

Bohr improved upon Rutherford's planetary model of the atom by stating that electrons exist in fixed energy states and emit photons only when they jump down energy states. The difference in energy between states becomes the energy of the emitted photon.



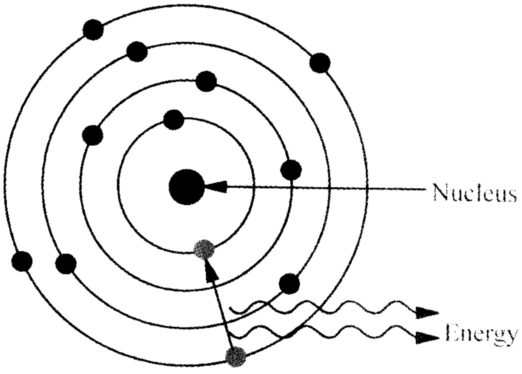
ELECTRON STATES

The phenomena of line emission and absorption spectra suggest that electrons can orbit nuclei only at certain energy levels. When electrons absorb energy, they jump to a higher energy level. This is called the "excited state" of the electron.



Atom

When an electron loses energy (emits light or a wavelength of light energy), it falls back down to the energy level it was in initially. This is called the "ground state" of the electron.

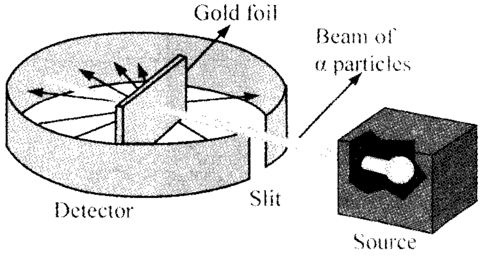


The Bohr and Rutherford models have been replaced by the electron cloud model, in which the probable locations of electrons of a given energy level are given by wave functions.

Practice

Use the following information to answer the next question.

In his famous alpha particle scattering experiment, Rutherford allowed a narrow beam of alpha (α) particles to strike a gold foil. He measured the angle of deflection of scattered alpha particles with a detector. He observed that most of the particles passed through the foil undeviated or suffered very little deflection.



1. Rutherford's alpha particle scattering experiment led to the confirmation of the
   A. position of electrons in an atom
   B. atomic model proposed by Thomson
   C. existence of a nucleus inside an atom
   D. existence of neutrons inside a nucleus

12.5.1.2 *describe the quantum mechanical model of the atom and the contributions of individuals to this model*

INNOVATORS IN ATOMIC STRUCTURE

The following chronology details the major developments in the atomic model in the twentieth century and the scientists who contributed to the developments.

1900: MAX PLANCK

The first major advance in understanding atomic structure was the idea that all energy is not continuous, but broken into small packets, called quanta. In 1900, Max Planck took his earlier research on blackbody radiation to a new level when his interpretations of experimental data led him to a new conclusion: that energy is not continuous. All light and radiation is emitted, transmitted, and absorbed in discrete amounts. The amount of energy contained in any wave of energy (E) can always be expressed as the frequency of that wave (Planck used the symbol v) multiplied by a constant.

1905: ALBERT EINSTEIN

A few years later, Albert Einstein theorized that light can behave as a particle or a wave (wave-particle duality of light). He explained the photoelectric effect by theorizing that light consists of photons, each of which constitutes a "quantum" of energy. Different wavelengths (λ) of light have different energies (E). The correlation between the energy of a wave and its wavelength is an inverse relationship, as shown in the equation

$$E = \frac{1}{\lambda}$$

1908–1911: ERNEST RUTHERFORD

In 1907, New Zealand physicist Ernest Rutherford assumed the position of the chair of physics at the University of Manchester. He started a laboratory and oversaw his students Hans Geiger and Ernest Marsden in performing the gold foil experiment. The results of the experiment led him to conclude that the nucleus is a relatively small and densely packed positive region of an atom.

1913: NIELS BOHR

One of Rutherford's students, Danish physicist Niels Bohr, expanded upon Rutherford's atomic theory by hypothesizing that electrons are in orbits that have specific amounts (quanta) of energy. An electron that falls to lower level emits energy, or a photon of light. An electron that is excited by absorbing energy expands its orbit, "jumping" to a higher orbital level.

1924: LOUIS DE BROGLIE

In his doctoral thesis published in 1924, French physicist Louis de Broglie hypothesized that matter such as electrons can have wave-like properties, thus having an assigned energy value (a quanta of energy). He extended his hypothesis to include every moving particle or object, which means that all moving matter has an associated wavelength.

1925: WOLFGANG PAULI

Wolfgang Pauli's exclusion principle established the foundation for an entirely new way of understanding solid matter. Pauli established that electrons are found in shells of different energy levels. He gave each shell a **quantum number** that expresses the distance of the shell from the nucleus. The **quantum state** of an electron is described by four numbers, determined by the quantum number, the shape of the orbit, the angular momentum, and the spin of the electron. The principle states that no two electrons within an atom or ion can share the same quantum state.

1926: FRIEDRICH HUND

After realizing that electron orbitals took specific shapes, German physicist Friedrich Hund derived a number of rules for how an electron "chooses" what type of orbital it will enter at a certain energy level. The rules are the basis for deriving the second number of an electron's quantum state, which describes whether it is in an s-, p-, or d-shaped orbital. An s orbital takes a spherical shape, a p orbital takes a barbell shape, and d orbitals have a more complex series of shapes. According to Pauli's exclusion principle, each orbital can hold only two electrons, one of each type of spin (the fourth number of the quantum state). Hund's rules state that electrons prefer to fill orbitals in the order of s, p, and then d. Within those subshells, electrons fill one orbital each before they start pairing up.

For example, a p orbital is barbell-shaped and will form three orbitals for each energy state, one along each of the *x*-, *y*-, and *z*-axes of the atom. If the atom acquires three electrons, each electron will move in a different orbital. No two electrons will occupy a single orbital in this case until a fourth is added. No electrons will begin to occupy d orbitals until all six p orbital spots have been filled.

1926: ERWIN SCHRÖDINGER

While at the University of Zürich in 1926, Erwin Schrödinger, an Austrian physicist, published a series of four papers that became the foundation for the field of physics known as wave mechanics. Wave mechanics provide the basis for generating the probability math used to locate and describe the behaviour of electrons. Schrödinger's work proved that electrons behave as both waves and particles.

1927: WERNER HEISENBERG

German physicist Werner Heisenberg is perhaps best known for the uncertainty principle that bears his name. The principle states that it is impossible to find the exact momentum and position of an electron at the same time (finding one disrupts the other). This proved Bohr's model to be incorrect and showed that electron shells are actually "smeared" in their shape, looking more like clouds than hard geometric forms.

SUMMARY OF THE QUANTUM MODEL

The quantum model uses four mathematical equations resulting in four numbers that describe each electron around an atom. Electrons exist in orbitals, which are areas where the probability of finding a given electron is high. Each electron has a specific energy, orientation in space, and spin.

Practice

2. No two electrons in an atom can have the same set of four quantum numbers, and no two electrons in an atom can occupy the same quantum state. This statement can be attributed to
   A. Hund's rule
   B. Avogadro's law
   C. the Pauli exclusion principle
   D. the Heisenberg uncertainty principle

3. Which physicist explained the dual nature of an electron?
   A. Bohr **B.** Thomson
   C. Avogadro **D.** Schrödinger

12.5.1.3 *list characteristics of the s, p, d, and f blocks of elements, and explain the relationship between position of elements in the periodic table, their properties, and their electron configurations*

QUANTUM NUMBERS

There are four quantum numbers for each electron around an atom. The numbers are derived from four mathematical equations, and each number describes a different property of the electron. The given table summarizes the meaning of each quantum number.

Meaning of Quantum Numbers

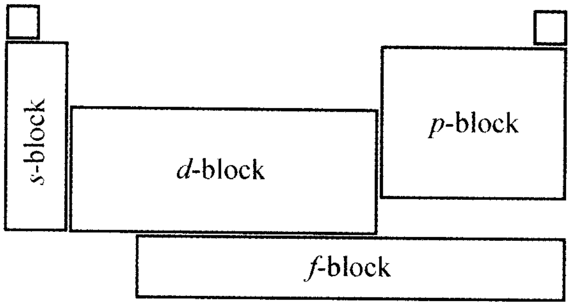
Quantum Number (QN)	**Energy Connection**	**Orbital Connection**	**Value**	**Connection to the Periodic Table**
Principal QN (n)	Main energy level	Shell	$n = 1 \rightarrow \infty$	Row
Secondary QN (l)	Energy sublevel	Orbital shape	$l = 0 \rightarrow n - 1$	Block
Magnetic QN (m_l)	Energy in magnetic field	Orbital orientation in space	$m_l = -l \rightarrow +l$ except $l = 0$	Columns (number of pairs of columns in each block)
Spin QN (m_s)	Individual electron energy difference	Electron spin	$m_s = -\frac{1}{2}, +\frac{1}{2}$	Individual column

The secondary quantum number (l) is expressed as a quantum number that represents different types of subshells within each main shell. These subshells are more commonly known as s, p, d, and f than by their numerical value. The given table summarizes the properties of each subshell.

Subshell Properties

Subshell	**Number of Orbitals in Subshell**	**Total Number of Electrons**
s	1	2
p	3	6
d	5	10
f	7	14

Quantum numbers are connected to the shape of the periodic table.



Practice

Use the following information to answer the next multipart question.

4. Identify the group of the periodic table that contains elements in which the last electron of the atom:

Part A

Open Response

enters but does not fill an s sublevel

Part B

Open Response

enters and fills an s sublevel

Part C

Open Response

is the first to enter a p sublevel

Part D

Open Response

is the second to last in a given p sublevel

Part E

Open Response

enters and fills a given p sublevel

Part F

Open Response

half fills a d sublevel

Open Response

5. Use the concept of s, p, d, and f orbitals to explain why noble gases are unreactive while alkali metals are quick to lose an electron.

12.5.1.4 *explain how the properties of a solid or liquid depend on the nature of the particles present and the types of forces between them*

INTERMOLECULAR FORCES

Intermolecular forces are the forces of attraction between molecules. Without intermolecular forces, all molecular substances would be in a gaseous state regardless of the temperatures, even at absolute zero.

THREE MAIN TYPES OF INTERMOLECULAR FORCES

1. **London forces**, also called dispersion forces or van der Waals forces, are present in all molecular substances. The greater the total number of electrons in a molecule, the greater the strength of the London forces. When comparing the strength of the London forces in two different molecular substances whose molecules have the same number of electrons, the substance whose molecules are shaped in such a way that they cannot approach each other closely is the one with the weaker London forces.
2. **Dipole-dipole forces** are present in polar molecular substances only. These forces are due to the attraction of the positive end, or pole, of one molecule to the negative end, or pole, of a neighbouring molecule.
3. **Hydrogen bonds** are extra-strong dipole-dipole forces present in molecular compounds with hydrogen (H) covalently bonded to fluorine (F), oxygen (O), or nitrogen (N). (You can use the mnemonic device "hydrogen FONding" to help you remember these elements.) Since F, O, and N have such high electronegativities compared with H, the covalent bond between H and the other atom leaves H with a very sparse electron cloud around it. This positively charged H is strongly attracted to the highly electronegative F, O, or N atom on neighbouring molecules, which have a denser electron cloud around those atoms, and are therefore more negatively charged.

RELATIVE BOND STRENGTH

London forces < Dipole-dipole < Hydrogen bonding

The stronger the intermolecular forces, the higher the melting and boiling points of a substance.

INTRAMOLECULAR FORCES

Other properties, such as electrical conductivity and malleability, depend on intramolecular forces, the forces within a molecule. Intramolecular forces include ionic bonds, covalent bonds, and metallic bonds. The more freely electrons can move about between atoms in a molecule, the higher the electrical conductivity of the substance and the more malleable the substance will be.

Practice

6. Covalent compounds have low melting points because they

A. are held together by weak van der Waals forces

B. are less exothermic than ionic compounds

C. have weaker bonds than ionic compounds

D. have definite shapes

Use the following information to answer the next multipart question.

7\.

Figure 1

X-H- - -:X

X = N, F, O

Figure 2

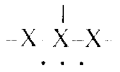


Figure 3

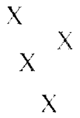


Figure 4

X Y X

Y X Y

Figure 5

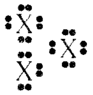
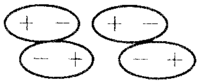


Figure 6



Part A

Open Response

Name the type of bond depicted in Figure 1.

Part B

Open Response

Name the type of bond depicted in Figure 2.

Part C

Open Response

Name the type of bond depicted in Figure 3.

Part D

Open Response

Name the type of bond depicted in Figure 4.

Part E

Open Response

Name the type of bond depicted in Figure 5.

Part F

Open Response

Name the type of bond depicted in Figure 6.

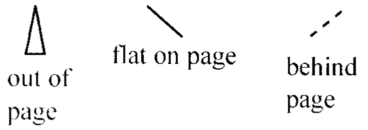
12.5.1.5 *explain how the Valence Shell Electron Pair Repulsion (VSEPR) model can be used to predict molecular shape*

VALENCE-SHELL ELECTRON PAIR REPULSION THEORY

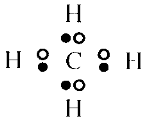
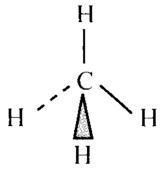
The valence-shell electron pair repulsion theory, or VSEPR theory, is used to predict the shape that the other atoms form around the central atom of a molecule. According to the theory,

- lone pairs and bonding pairs around a central atom move as far apart as possible to minimize repulsion
- lone-pair electrons repel with greater force than bonding-pair electrons, creating wider angles around lone electron pairs
- multiple bonding pairs are treated as though they were single bonding pairs

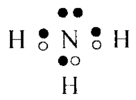
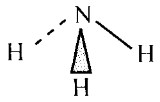
Key:



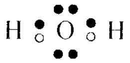
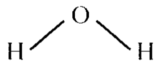
METHANE

Formula	CH_4
Lewis Diagram	
Lone Pairs	0
Bonding Pairs	4
Shape Name	Tetrahedral bond Angle = 90°
Shape Diagram	

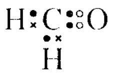
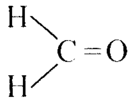
AMMONIA

Formula	NH_3
Lewis Diagram	
Lone Pairs	1
Bonding Pairs	3
Shape Name	Trigonal pyramidal bond Angle = 109°
Shape Diagram	

WATER

Formula	H_2O
Lewis Diagram	
Lone Pairs	2
Bonding Pairs	2
Shape Name	V-shaped bond Angle = 109°
Shape Diagram	

METHANAL

Formula	CH_2O
Lewis Diagram	
Lone Pairs	0
Bonding Pairs	3
Shape Name	Trigonal planar bond Angle = 120°
Shape Diagram	

HYDROGEN CYANIDE

Formula	HCN
Lewis Diagram	
Lone Pairs	0
Bonding Pairs	2
Shape Name	Linear bond Angle = 180°
Shape Diagram	$H-C\equiv N$

Practice

8. Which of the following compounds has two lone pairs and two bond pairs of electrons?

A. NH_3 **B.** BF_3

C. H_2O **D.** CO_2

Open Response

9. Methane (CH_4) and ammonia (NH_3) have the same number of electrons around their central atom, but they do not have the same shape. Determine the shape of each molecule and explain their difference in shape.

12.5.2.1 *use appropriate scientific vocabulary to communicate ideas related to structure and bonding*

VOCABULARY FOR MOLECULAR STRUCTURE AND BONDING

For scientists to have meaningful discussions about the chemical properties of elements, atoms, compounds, and molecules, the language and terms they use must be clear and precise.

ELECTRON STRUCTURES

Electrons in atoms or ions are found in energy levels called **shells**. In order of increasing energy state, the first electron shell contains two electrons, the second contains eight, the third contains 18, and the fourth contains 32. (Higher energy states also exist.) The difference in energy between one electron shell and another is different for each element. The first electron shell is numbered 1, the second 2, and so on, which is the first value in an electron's quantum number.

Each electron shell contains subshells, or **orbitals**, of which there are several types. In order of ascending energy value, the first four orbital forms are s, p, d, and f, and each has a different three-dimensional shape. Each orbital can hold two electrons, one of each type of spin. The orbital shapes are numbered starting at 0, which forms the second value in an electron's quantum number.

The s orbital is a single spheroids within an energy level and can hold two electrons. The p orbitals have a barbell shape and extend along the *x* -, *y*-, and *z*-axes of an atom, with the nucleus as the vertex. The p orbitals can therefore hold six electrons per energy level. The d and f orbitals have variable shapes and can hold 10 and 14 electrons, respectively.

According to Hund's rules, electrons have a specific pattern in which they orbit a nucleus. An orbital will not contain more than one electron until all of the orbitals of the same type already contain at least one electron. For p orbitals, this means that the first three electrons to orbit the atom will each occupy its own p orbital before they will begin to pair up. The Aufbau principle states that orbitals of lower energy levels are filled before higher-energy orbitals, so s orbitals are filled before p orbitals, p orbitals are filled before d orbitals, and so on. The Pauli exclusion principle states that no two electrons in the same orbital can have the same quantum state, so they must have opposite spins.

ELECTRONS AND ENERGY

When electrons emit or absorb energy, they can shift from one energy level to another. Energy is transmitted in discrete packets called **quanta**. Each quantum has both a wave and a particle nature. The wave nature of a quantum is directly related to its wavelength.

A quantum packet of light energy is known as a **photon**. When a range of photons are projected at an atom, electrons will absorb the photons that have a wavelength equal to the energy needed to make the electrons jump to a higher shell level. Each element therefore has its own signature **absorption spectrum**, as the remaining photons form a spectrum with black bars showing the wavelengths that did not pass through the atom. Conversely, when electrons jump down to a lower shell level, they emit the photons with wavelengths that correspond to the energy lost, creating an **emission spectrum**.

BONDING

The noble gases are the most stable elements because their outer electron shell structures have a full complement of electrons. The other elements are less stable because they are missing one or more electrons from their outer shell, so their atoms tend to share electrons between orbitals. This sharing forms **covalent bonds** between the atoms.

An atom that achieves a stable electron structure and consequently has more electrons than protons becomes a negative ion. In the same way, an atom that forms a stable outer electron shell by losing electrons becomes a positive ion. The tendency of an atom to gain or lose electrons is called its **valency**, and electrons that tend to be lost are known as **valence electrons**. Positive and negative ions attract and bond with each other to reach a neutral state of electronegativity, forming **ionic bonds**.

Atoms or ions that have bonded to form molecules may have either an even or an uneven distribution of electrons around the group of nuclei. Molecules that generally have more electrons present at one end than the other are referred to as **polar molecules**. Polar molecules can form **dipole-dipole bonds** with one another, as the positive end of one molecule is attracted to the more electron-dense negative end of another. When hydrogen bonds covalently with the highly electronegative element fluorine, oxygen, or nitrogen, the hydrogen atom loses most of its ability to attract electrons. This causes it to form a positive end on the molecule. The positive end is strongly attracted to the fluorine, oxygen, or nitrogen atom of a neighbouring molecule, forming a **hydrogen bond**—an extra-strong dipole-dipole bond. Hydrogen bonding can occur both intra- and intermolecularly.

Practice

10. The area around the nucleus of an atom where the probability of finding an electron is the highest is called

A. a dipole

B. an azide sphere

C. an orbital

D. a quantum level

12.5.2.2 *write electron configurations for elements in the periodic table, using the Pauli exclusion principle and Hund's rule*

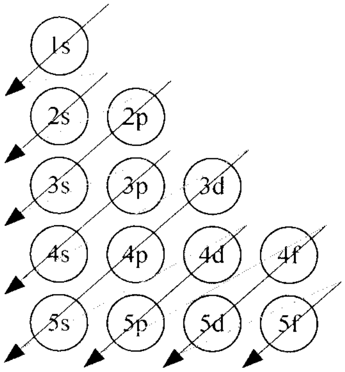
ELECTRON CONFIGURATION

ATOMS

Electron configuration is the arrangement of electrons around an atom, according to the quantum model. The standard notation used to communicate the electron configuration of an atom is written as nx^y, where n is the shell number (1, 2, 3, etc.), x is the subshell (s, p, d, f, etc.), and y indicates the number of electrons in the subshell.

Several principles have been developed to explain the way in which electrons fill up an atom's subshells:

1. **Aufbau principle**—In the ground state, electrons fill lower energy subshells first.
   The given figure illustrates the order in which the subshells fill up.



2. **Hund's rule**—Electrons do not pair up unless they must. There must be one electron in each orbital before electrons will begin to pair up.
3. **Pauli exclusion principle**—No two electrons can have the same four quantum numbers, so electrons in the same orbital must have different spin numbers. Electron spin is represented by ↑ and ↓.

The electron configuration of any element in the periodic table can be determined by using these principles. For example, the electron configuration of oxygen is $1s^2 2s^2 2p^4$.

A shorthand form is used for elements with long electron configurations. The nearest preceding noble gas is written, and the electron configuration is written as it occurs after the noble gas.
For example, the electron configuration for bromine is $[Ar]4s^2 3d^{10} 4p^5$. Notice that the orbitals are not listed by increasing principle quantum number. This is in accordance with the Aufbau principle which, for example, states that the 4s orbital is filled before the 3d orbital, as shown in the previous diagram.

IONS

In representing the electron configuration of an ion, the electrons are taken from the orbital with the highest principle quantum number (n). This means that when expressing the electron configuration of an ionizing metal, the orbitals are listed in order of the increasing principle quantum number, not in conventional Aufbau principle notation.
For example, the electron configuration for a gallium atom is $[Ar]3d^{10} 4s^2 4p^1$, and the electron configuration for Ga^{3+} is $[Ar]3d^{10}$.

Practice

11. The electronic configuration of which of the following is represented as $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$?

A. Co **B.** Zn

C. Co^{2+} **D.** Zn^{2+}

Open Response

12. Use shorthand notation to write the electron configuration for tin and its two most common ions.

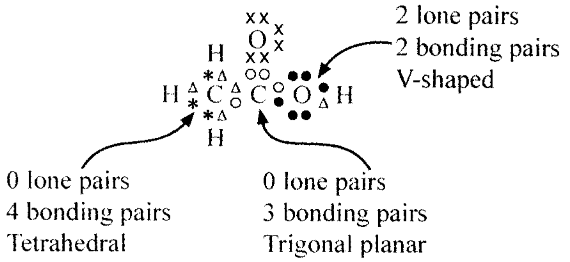
12.5.2.3 *predict molecular shape for simple molecules and ions, using the VSEPR model*

USING VSEPR THEORY TO PREDICT MOLECULAR SHAPE

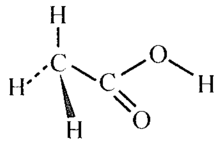
The principles of the VSEPR theory can be used to predict the shape of a molecule even when it has more than one central atom. This is done by examining the VSEPR arrangement around each central atom individually.

The given Lewis and shape diagrams illustrate the molecular shape of CH_3COOH.

Lewis Diagram of CH_3OOH



Shape Diagram of CH_3OOH



Making models of molecules is a good way to equate the diagrams with the three-dimensional shape of the molecules.

Practice

13. According to the VSEPR theory, what shape is a BF_3 molecule?

A. Pyramidal

B. Tetrahedral

C. Square planar

D. Trigonal planar

Use the following information to answer the next multipart question.

14. Using the VSEPR theory, draw and name the shape of the following molecules:

Part A

Open Response

$SiCl_4$

Part B

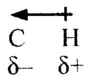
Open Response

NO_2^-

12.5.2.4 *predict the polarity of various substances, using molecular shape and the electronegativity values of the elements of the substances*

MOLECULAR SHAPE, ELECTRONEGATIVITY, AND POLARITY

When two atoms do not share electrons equally, the bond between them is called a **polar covalent bond**. The atom with higher electronegativity is partially negative, $\delta-$, while the one with lower electronegativity is partially positive, $\delta+$. For example, in the bond between carbon and hydrogen, carbon has the higher electronegativity and is $\delta-$, while hydrogen has the lower electronegativity and is $\delta+$. The given figure illustrates how to represent this bond.



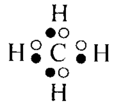
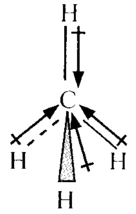
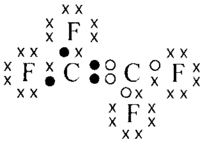
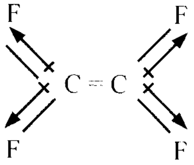
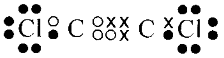
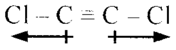
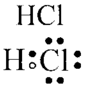
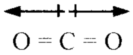
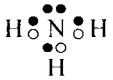
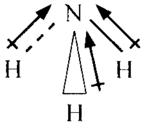
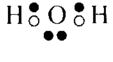
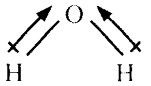
A polar covalent bond

The arrow represents a bond dipole, where the arrowhead is the negative end and the tail is the positive end of the polar covalent bond.

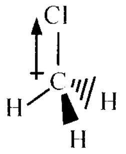
Just as bonds can be polar or non-polar, so too can molecules. A molecule's polarity affects a number of its physical properties, including its solubility, melting point, and boiling point.

To find out whether a molecule is polar or not, predicts its shape and the polarity of its bonds. By considering both the size and direction of the dipoles, you can determine whether the polarities of the bonds cancel each other out or produce a net dipole.

Examples of Polar and Non-polar Molecules

Formula and Lewis Diagram	**Shape Diagram with Bond Dipoles**	**Polar or Non-polar**
CH_4 	 Tetrahedral	Non-polar
C_2F_4 	 Trigonal planar	Non-polar
C_2Cl_2 	 Linear	Non-polar
Br_2 	Br – Br Linear	Non-polar
HCl 	 Linear	Polar
CO_2 	 Linear	Non-polar
NH_3 	 Trigonal pyramidal	Polar
H_2O 	 V-shaped	Polar

Molecules that are symmetrical in shape (e.g., tetrahedral, linear, trigonal planar) can have polar bonds, but the molecule as a whole is not polar. The symmetrical shape of the molecule causes the dipoles to cancel each other out, making the molecule non-polar (e.g., CH_4, C_2F_2, and CO_2). If the dipoles are not equal, the molecule is polar, as can be seen in the following example of chloromethane.



Chloromethane

Practice

Use the following information to answer the next question.

The product of the magnitude of the positive charge (q) and the distance (d) between the centres of the negative and positive charges within a molecule is called the dipole moment. It is represented mathematically by the symbol μ and is calculated using the given formula.
$\mu = q \times d$

15. Which of the following compounds has the highest dipole moment?

A. HI **B.** HF **C.** HCl **D.** HBr

Open Response

16. $C_2H_2F_2$, which has a carbon-carbon double bond, can form both polar and non-polar structural isomers. Use a diagram to explain how this is possible.

Open Response

17. Use the VSEPR theory to draw the shape of each of the following molecules. Name their shape and determine if they are polar or non-polar.

Molecule	VSEPR 3-D Shape	Name of Shape	Polar? (Y/N)
HC_2F			
SCl_2			
$SiCl_4$			
$N_2H_3F_2$			

12.5.2.5 *predict the type of solid (ionic, molecular, covalent network, or metallic) formed by a substance, and describe its properties*

12.5.2.6 *conduct experiments to observe and analyse the physical properties of different substances, and to determine the type of bonding present*

TYPES AND PROPERTIES OF SOLIDS

There are four main classes of crystalline solids: metallic, molecular, covalent network, and ionic. The types of bonding within the molecules and the intermolecular forces present in these solids influence their physical and chemical properties. Knowing the properties of each type of solid is useful in identifying solids through experimental analysis.

METALLIC SOLIDS

DESCRIPTION

Positive nuclei are embedded in a sea of electrons. The electrons hold the nuclei together with very strong intermolecular forces.

CHARACTERISTICS

- conduct electricity
- malleable
- melting and boiling points vary significantly between solids in this class

EXAMPLES

- Fe
- steel
- Pb
- Cu

MOLECULAR SOLIDS

DESCRIPTION

Neutral atoms are covalently bonded, and the molecules are held together by weak intermolecular forces (London forces, dipole-dipole forces, and hydrogen bonding in some cases).

CHARACTERISTICS

- low melting and boiling points
- poor conductors of electricity

EXAMPLES

- H_2O (ice)
- CO_2 (dry ice)
- I_2

COVALENT-NETWORK SOLIDS

DESCRIPTION

Atoms are in strong directional covalent bonds. This crystal is like one big macromolecule because of the arrangement of molecules into chains and webs.

CHARACTERISTICS

- very hard
- high melting and boiling points (higher than ionic solids)
- non-conductive
- not malleable

Example

- $C_{(graphite)}$
- $C_{(diamond)}$
- SiO_2
- SiC

IONIC SOLIDS

DESCRIPTION

Ions are attracted to each other through electrostatic attraction, forming a crystal lattice.

CHARACTERISTICS

- strong and brittle
- conduct in molten state or if dissolved in water
- high melting and boiling points

Example

- NaCl
- $CaCO_2$

Practice

18. Which of the following types of bond exist in solids with the **best** thermal and electrical conductivity?

A. Ionic　　**B.** Metallic

C. Covalent　　**D.** Van der Waals

19. If tested experimentally, which of the following substances would be observed to have the highest melting point?

A. SCl_2　　**B.** NaCl

C. SiO_2　　**D.** NH_3

Open Response

Use the following information to answer the next question.

The boiling point of a molecular compound is a measure of the type and extent of bonding between the molecules of that compound. In an experiment, a chemist heated four unidentified organic samples to their boiling point. The given table provides the data gathered from the experiment.

Sample	Boiling Point (°C)	Identity
1	47.9	
2	–2.5	
3	97.2	
4	–42.1	

20. The chemist knows that each of the samples is either $CH_3CH_2CH_3$, CH_3CH_2F, $CH_3CH_2NH_2$, or CH_3CH_2OH. Using the experimental data, determine the identity of each sample and explain your answer.

12.5.3.1 *describe some applications of principles relating to atomic and molecular structure in analytical chemistry and medical diagnosis*

12.5.3.3 *describe advances in Canadian research on atomic and molecular theory*

12.5.3.2 *describe some specialized new materials that have been created on the basis of the findings of research on the structure of matter, chemical bonding, and other properties of matter*

APPLICATIONS AND DEVELOPMENTS IN ATOMIC AND MOLECULAR STRUCTURE

An understanding of how molecular and atomic structure affect the properties of a substance was crucial in the development of many commonly used materials, such as bulletproof vests, tires, super glue, and superconductors (used in transport and magnetic resonance imaging). Scientists study molecular structure using X-ray crystallography and infrared spectroscopy. These tools are used by pharmaceutical companies to develop new drugs or to test medications to ensure that the correct medication and dose are administered to patients. They are also used in forensic analysis to identify chemicals and in industry to test the purity of a chemical product. Technologies such as nuclear medicine (both diagnostic and chemical treatment) would not be possible without a thorough understanding of the molecular structure and properties of the chemical species involved.

The study of atomic theory is far from over. In fact, researchers, including many Canadian scientists, continue to pursue a more comprehensive model of atoms and the molecules they form. As scientists learn more about the atomic model, they will be able to develop more applications and technologies that rely on the understanding of atomic structures and properties.

Practice

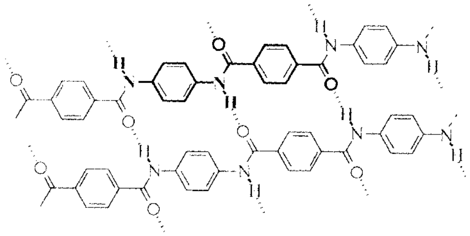
Open Response

21. In the field of nuclear medicine, radioactive material is used to diagnose and treat various illnesses that occur within the human body. Use the Internet or other sources to research the following questions regarding nuclear medicine, and report your findings in a brief essay.

- What are four common uses for nuclear medicine?
- What are four chemical species used in nuclear medicine, and how are they administered to patients?
- How does nuclear imaging work?
- What types of treatments involve radioisotopes, and how do they work?
- What are two benefits and two risks of using nuclear medicine to diagnose and treat human illness?

Use the following information to answer the next multipart question.

22. Kevlar is a lightweight and extremely strong polymer used to fabricate products such as tires, racing sails, and bulletproof vests. The polymer's structure is depicted in the given figure.



Part A

Open Response

Name and explain at least three of the properties of Kevlar that make it a good material to use for bulletproof vests.

Part B

Open Response

What is one of Kevlar's major weaknesses as a bulletproof material, and what do manufacturers of bulletproof vests do to compensate for this weakness?

Open Response

23. Although the spotlight has often been on European scientists for their advances in atomic theory, many Canadian scientists have contributed to atomic and molecular theory through their research and new technological developments. Using the Internet and other sources, research the contributions of either Richard Bader or R. J. LeRoy (the LeRoy radius), and report your findings in a short paragraph. Include the following information about the scientist you choose:

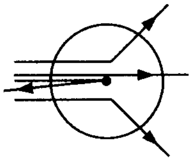
- What is the focus of his research?
- What are the possible practical implications of that research?
- How long has he been working on his research?

SOLUTIONS–STRUCTURE AND PROPERTIES

1. C
2. C
3. D
4. Part A- **OR**
Part B- **OR**
Part C- **OR**
Part D- **OR**
Part E- **OR**
Part F- **OR**
5. OR
6. A
7. Part A- **OR**
Part B- **OR**
Part C- **OR**
Part D- **OR**
Part E- **OR**
Part F- **OR**
8. C
9. OR
10. C
11. D
12. OR
13. D
14. Part A- **OR**
Part B- **OR**
15. B
16. OR
17. OR
18. B
19. C
20. OR
21. OR
22. Part A- **OR**
Part B- **OR**
23. OR

1. C

The pattern of the alpha particles scattered by the gold foil shows that the mass of an atom is not distributed uniformly, but is concentrated mainly at a positively charged centre.



Only a few alpha particles bounced back after striking this centre. This shows the existence of a dense nucleus inside an atom.

2. C

The Pauli exclusion principle states that no two electrons in an atom or an ion can have the same quantum state, as defined by the four quantum numbers.

3. D

Schrödinger postulated a wave equation that incorporated both the wave nature and the particle nature of an electron.

4. Part A – Open Response

Group IA

Part B – Open Response

Group IIA

Part C – Open Response

Group IIIA

Part D – Open Response

Group VIIA

Part E – Open Response

Group VIIIA

Part F – Open Response

Group VIIB

5. Open Response

Noble gases have full p orbitals, making them stable because they do not need to lose or gain any electrons. Alkali metals have one electron in the s orbital, and losing that electron would leave an alkali atom with a full p orbital, which is more stable. Therefore, alkali metals lose one electron easily from their outer s orbital.

6. A

There is a small electronegative difference present between the atoms in covalent compounds, so groups of those type of molecules are held together by weak van der Waals forces. Because their intermolecular forces are weak, covalent compounds have low melting points.

7. Part A – Open Response

Hydrogen bonding

Part B – Open Response

Covalent network

Part C – Open Response

Metallic

Part D – Open Response

Ionic

Part E – Open Response

London forces / dispersion forces / van der Waals forces

Part F – Open Response

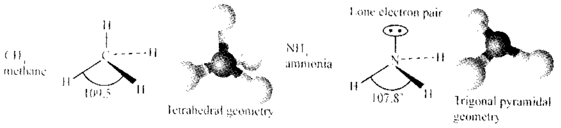
Dipole-dipole

8. C

The compound H_2O has two lone pairs and two bond pairs of electrons. Oxygen is the central atom and has six valence electrons. Two hydrogen atoms mutually share two electrons, while the remaining four form two lone pairs.

Ammonia (NH_3) has one lone pair and three bond pairs of electrons. The compound BF_3 has no lone pair of electrons, though it has three bond pairs of electrons. Carbon dioxide (CO_2) has no lone pair of electrons, though it has two bond pairs of electrons.

9. Open Response



Methane is tetrahedral while ammonia is trigonal pyramidal. The difference in shape between the two molecules is caused by ammonia's lone pair of electrons (electrons that are not bonding). The pair repels bonding electrons, causing them to move closer together in ammonia than they are in methane. The bond angle between the hydrogen atoms in ammonia (107.8°) is therefore smaller than the bond angle between the hydrogen atoms in methane (109.5°).

10. C

An orbital is defined as the area around the nucleus where an electron is most likely to be located.

11. D

The atomic number of the given element as calculated by its electronic configuration is
$2 + 2 + 6 + 2 + 6 + 10 = 28$
The element with the atomic number of 28 is nickel, but it is not one of the possible answers.
The atomic number of zinc is 30, and its electronic configuration is represented as
$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$. If zinc loses two electrons, it becomes Zn^{2+}, which has an electronic configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$.

12. Open Response

Sn: $[Kr]4d^{10}5s^2 5p^2$

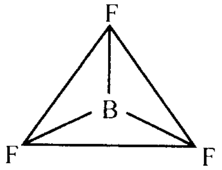
Sn^{2+}: $[Kr]4d^{10}5s^2$

Sn^{4+}: $[Kr]\ 4d^{10}$

Electrons are taken first from the suborbitals that have the highest principal quantum number, which in this case is $n = 5$. The outermost, or most energized suborbitals with an identical principle quantum number are also the most likely to lose valence electrons. In order, d orbitals will lose electrons before p orbitals, which in turn are also more likely to lose electrons than s orbitals.

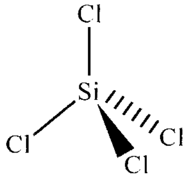
13. D

In a BF_3 molecule, the central metal atom, boron, has three valence electrons that are shared equally by the three fluorine atoms. The bond angle FBF is 120°. According to the VSEPR theory, the shape of the BF_3 molecule is trigonal planar.



14. Part A – Open Response

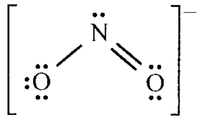
$SiCl_4$



Tetrahedral

Part B – Open Response

NO_2^-



V-shaped

15. B

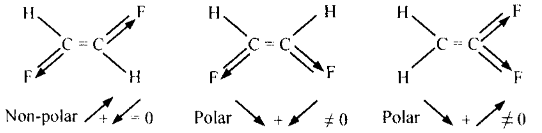
The dipole moment of a compound depends on the electronegativity difference between the constituent atoms. The greater the electronegativity difference, the greater the dipole moment of the compound. All four alternatives are hydrogen compounds, so the given table shows the electronegativities of the halogens.
The compound containing the halogen with the greatest electronegativity is the compound with the highest dipole moment.

Halogen	Electronegativity
I	2.5
F	4.0
Cl	3.0
Br	2.8

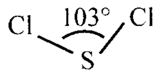
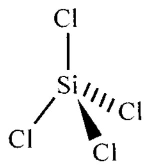
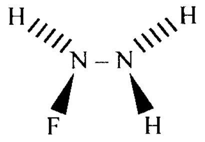
Fluorine is the most electronegative element of the given halogens, so HF has the highest dipole moment of the given compounds.

16. Open Response

Shapes of Difluoroethene



17. Open Response

Molecule	VSEPR 3-D Shape	Name of Shape	Polar? (Y/N)
HC_2F	H – C ≡ C – F	Linear	Y
SCl_2		V-shaped	Y
$SiCl_4$		Tetrahedral	N
N_2H_3F		Trigonal planar	Y

18. B

Metallic bonds exist between metals, and are usually the best conductors of heat and electricity. The valence electrons of individual atoms that make up the metal are not closely associated, therefore electrons can move freely between the atoms that make up the metal.

Because electrons can move freely, an electrical current and heat can flow through the metal easily.

19. C

SiO_2 is a covalent network compound that consists of a lattice of strong covalently bonded silicon and oxygen atoms. The other given compounds have weaker intermolecular forces, so less energy is required to break those intermolecular forces. The stronger the bonds in a substance, the higher its melting point.

20. Open Response

Sample	Boiling Point (°C)	Identity
1	47.9	$CH_3CH_2NH_2$
2	−2.5	CH_3CH_2F
3	97.2	CH_3CH_2OH
4	−42.1	$CH_3CH_2CH_3$

Propane, $CH_3CH_2CH_3$, has the fewest intermolecular bonds (only London forces), so it has the lowest boiling point compared with the other compounds. Fluoroethane, CH_3CH_2F, has both London forces and dipole-dipole intermolecular forces; it has the second-lowest boiling point of the given compounds. The boiling point of ethanol, CH_3CH_2OH, exceeds that of ethanamine, $CH_3CH_2NH_2$, because —OH functional groups tend to have stronger hydrogen bonds than N—H functional groups: the —OH bond is more polar. Both ethanol and ethanamine have boiling points higher than fluoroethane because hydrogen bonds are stronger than London forces and dipole-dipole forces.

21. Open Response

Nuclear medicine involves the use of radioactive pharmaceuticals (radiopharmaceuticals) to diagnosis and treat certain diseases. Some common uses for nuclear medicine are bone scans (for tumours or injury), heart scans, various tumour scans (ovarian, prostate, colorectal, lung), and three-dimensional diagnostic imaging of the entire body.

Radiopharmaceuticals can be given intravenously, inhaled, or ingested with food. Iodine-123 and 131, thallium-201, and gallium-67 are given through intravenous injection, while xenon-133 and krypton-81 are examples of gaseous radioisoptoes administered by inhalation. The radioisotope is chemically linked to another chemical (pharmaceutical) called a tracer that tends to be absorbed more readily in certain areas of the body. Once the radiotracer reaches the desired location in the body, the radioisotope emits radiation (gamma rays) that is detected by a gamma camera or positron emission tomography scanner to create a three-dimensional image of the area or organ. Usually, the diseased area of an organ will attract either more or less of the tracer than the healthy areas, enabling doctors to locate and determine the extent of the diseased tissue.

The radioisotopes can also be used to treat diseases. Isotopes with higher levels of radiation (such as I-131) are used to treat thyroid gland diseases, for example, because the radiation kills the abnormal cells.

There are both benefits and risks associated with nuclear medicine. Nuclear imaging provides a more accurate and informative image of diseased tissue than other diagnostic methods, which can improve diagnosis and treatment.

It is also less invasive than surgical examination and can often be used to identify organ abnormalities in the early stages of a disease. As with any medical procedure, there are also risks, such as exposure to low doses of radiation, which is of particular concern for pregnant or nursing women. Some patients have had allergic reactions to the pharmaceuticals used, and the intravenous injection of radiotracers can be painful. Nuclear imaging can also be time-consuming, taking several hours to perform. Overall, nuclear medicine has been a widely used and accepted form of diagnosis and treatment for more than 50 years, with no known long-term side effects.

22. Part A – Open Response

Kevlar, or poly-para-phenylene terephthalamide, is composed of many long polymer chains with many intermolecular hydrogen bonds between adjacent chains, which make the material extremely strong. The benzene rings stack on top of each other, causing an intermolecular overlapping of pi (π) bonds, which adds to the strength of the intermolecular attraction between Kevlar chains. Kevlar is produced in such a way that the polymers orient in long, straight chains, which maximizes intermolecular attraction and creates a strong but not brittle crystalline structure. Finally, the polymer contains no heavy salts, making the material even more lightweight than nylon.

Part B – Open Response

Kevlar decomposes in ultraviolet light (i.e., the sun). To minimize the material's exposure to UV light, manufacturers cover the Kevlar in bulletproof vests with another material.

23. Open Response

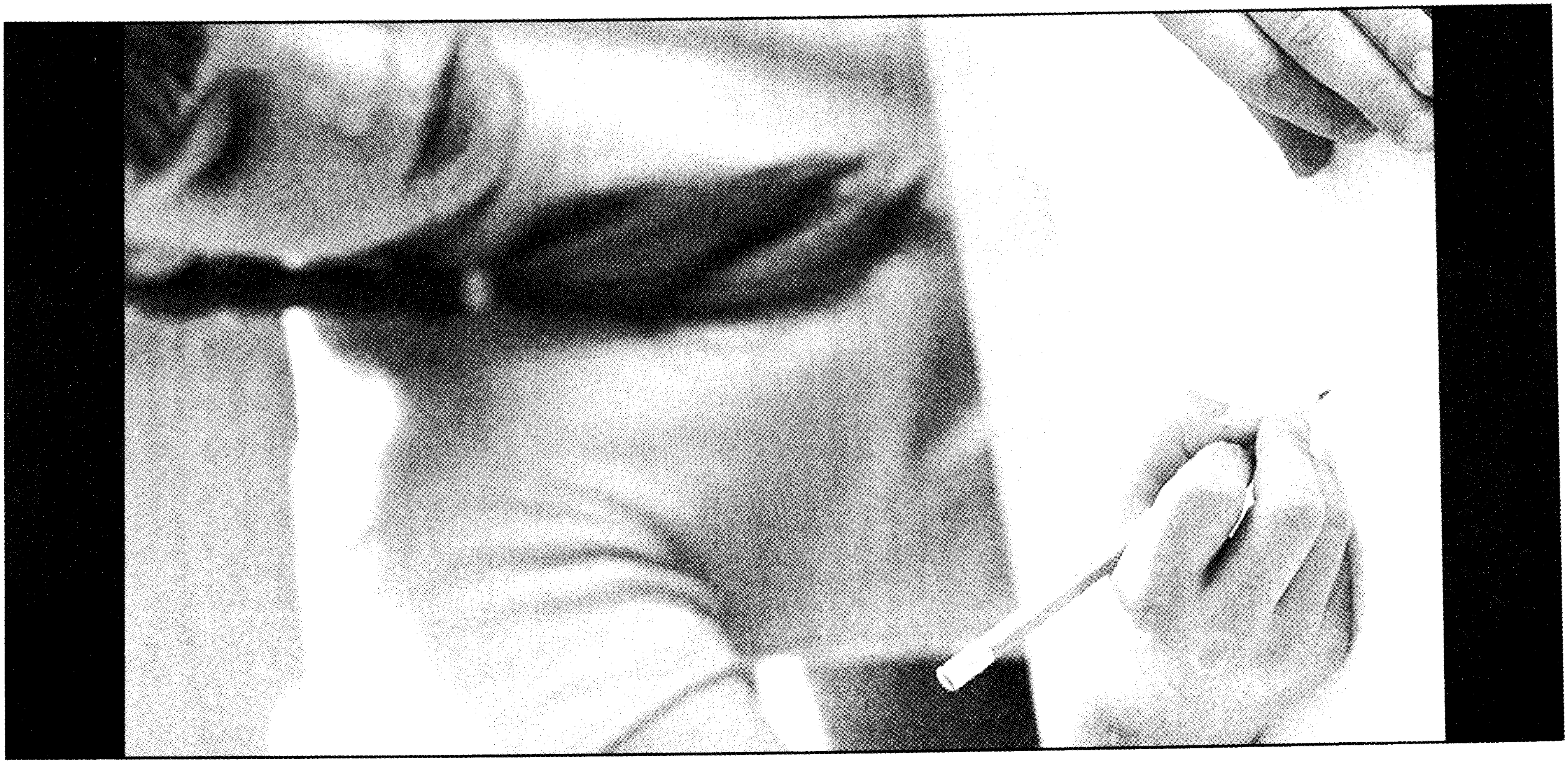
Richard Bader

Richard Bader's theory is called the Quantum Theory of Atoms in Molecules, which states that bonds are not overlapping electron orbitals; rather, areas of electron density exist between atomic nuclei in a molecule and can be mathematically derived. The physical properties of a molecule can be determined by listing and defining the properties of constituent atoms. A practical implication of this research is that computers rather than laboratory experiments could one day be used to determine molecular properties. Bader has worked in chemistry for 55 years and has been working on electron density for approximately 20 years.

R. J. LeRoy

R. J. LeRoy derived the LeRoy radius, which is the radius of atoms in a molecule that are about to dissociate. The atoms are at a distance from each other where intermolecular (physical), not intramolecular (chemical), forces are dominant. The LeRoy radius is useful for determining bond dissociation energies, which helps in predicting the thermodynamic properties of molecules and the equilibrium dissociation constants for high-temperature reactions. LeRoy has worked in chemistry for 48 years and has been working on the LeRoy radius for eight years.

Self Test



1. Which of the following statements does **not** support Bohr's atomic model?

   A. Electrons only occupy an orbit or energy level in which the angular momentum of the electron is an integral multiple of $\frac{h}{2\pi}$, where h = Plank's constant.

   B. If sufficient energy is available, an electron will absorb energy and move to a higher energy level.

   C. An atom consists of two distinct parts: the nucleus and a single electron shell around it.

   D. Electrons always move within certain orbits that have specific energies.

Open Response

2. The model of the atom has been modified and experimentally tested for well over 100 years. Many scientists have contributed to the ever-changing understanding of the atomic model. In a brief paragraph, identify and explain Max Planck's and Albert Einstein's major contributions to the development of the quantum model of the atom, using the given equation and terms in your explanation.
   - $E = h \times v$
   - photoelectric effect
   - quantum or quanta
   - photon

____

____

____

____

____

3. The block in the periodic table that contains elements with valence configurations between s^2p^1 and s^2p^6 is the

   A. d-block **B.** f-block

   C. p-block **D.** s-block

Use the following information to answer the next question.

The air in the vicinity of areas of geothermal activity, such as Yellowstone National Park, have a characteristic "rotten egg" aroma. The smell is due to the toxic gas hydrogen sulfide (H_2S).

Water (H_2O) is a liquid at standard atmospheric temperature and pressure, but hydrogen sulfide, hydrogen selenide (H_2Se), and hydrogen telluride (H_2Te) are all gases under the same conditions.

4. Water has a higher boiling point than the other given compounds because of its

   A. hydrogen bonds **B.** covalent bonds

   C. metallic bonds **D.** ionic bonds

Open Response

5. The molecular shape of sulfur dioxide, SO_2, is V-shaped, but there are three different forms in which the molecule can exist in this shape. Using electron dot diagrams, draw the three forms in which SO_2 can exist, and explain why this molecule has three simultaneous structures.

6. The ground state electronic configuration of strontium is
   - **A.** [Kr] $5s^1$
   - **B.** [Kr]$5s^2$
   - **C.** [Kr]$4d^{15}s^2$
   - **D.** [Kr]$4d^{25}s^2$

Open Response

7. What is the electron configuration of the two most common ions for thallium, atomic number 81?

Use the following information to answer the next multipart question.

8. The International Union of Pure and Applied Chemistry nomenclature rules apply systematically to chemical compounds. Some other common names for compounds exist, and some are recognized by IUPAC as exceptions to the systematic naming conventions. For example, the formal name of C_2H_4 is ethene, but many people refer to it by its common name of ethylene.

Part A

Open Response

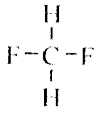
Using the VSEPR theory, draw the shape of C_2H_4 and name the shape around the central atom(s).

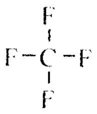
Part B

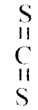
Open Response

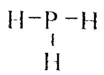
Using the VSEPR theory, draw the shape of CH_3OH and name the shape around the central atom(s).

Use the following information to answer the next question.


I
$F-\overset{H}{\underset{H}{C}}-F$


II
$F-\overset{F}{\underset{F}{C}}-F$


III
$S=C=S$


IV
$H-\underset{H}{P}-H$

9. Which of the given molecules is polar?
A. I **B.** II **C.** III **D.** IV

Numerical Response

Use the following information to answer the next question.

	Six Molecules
1	*ZX* linear
2	ZX_2 linear
3	ZX_2 bent
4	ZX_4 tetrahedral
5	ZX_3 trigonal pyramidal
6	ZX_5 square pyramidal

10. Assuming the difference in electronegativity between *X* and *Z* is greater than 0.5, which of the following molecules can have a nonzero dipole moment? ____ ____ ____ ____

Use the following information to answer the next question.

The two types of intermolecular forces that are demonstrated by H_2S are _*i*_ and _*ii*_.

11. The information in which of the following tables completes the given statement?

A.

i	*ii*
London forces	hydrogen bonding

B.

i	*ii*
dipole-dipole forces	London forces

C.

i	*ii*
hydrogen bonding	dipole-dipole forces

D.

i	*ii*
dipole-dipole forces	ionic bonds

Open Response

Use the following information to answer the next question.

A student is testing the compounds in the given table for their solubility in water and isopropanol.

Compound	Solubility in Water	Solubility in Isopropanol
NaCl		
$C_{(graphite)}$		
Wax Shavings		

12. Hypothesize the solubility of each compound in both water and isopropanol. Include the rationale for your answer.

SOLUTIONS

1. C	5. OR	Part B- OR	12. OR
2. OR	6. B	9. A	
3. C	7. OR	10. 1 3 5 6	
4. A	8. Part A- OR	11. B	

1. C

The theory that an atom consists of only two distinct parts was proposed by Rutherford based on the results of experiments in which he observed that alpha particles were scattered by thin metal foil. It supports Rutherford's nuclear model of the atom, but not Bohr's atomic model, which described different electron orbitals and states of excitation.

2. Open Response

Max Planck developed the idea that energy is not continuous, but rather increases in levels, or quanta. His equation $E = hv$ shows that the energy (E) of a vibrating molecule is a multiple of the vibration frequency (v). The multiple is Planck's constant (h), which is

$6.626 \times 10^{-34}\ \text{J} \cdot \text{s}$

This idea of quantized energy was applied to light by Alberta Einstein, who used Planck's idea to explain the photoelectric effect—electrons emitted from a heated piece of metal. Einstein theorized that electrons are ejected from the surface of the metal only when the light particle, or photon, contains the proper amount of threshold energy. The energy is passed to the electron, which causes it to break free of the attractive forces of the metal.

Plank's and Einstein's findings led to the ideas that electron orbitals have set sizes and energies and that light can behave as a particle or a wave (wave–particle duality of light), since a photon is absorbed and ceases to exist when it hits an electron.

3. C

The elements in the p-block have valence configurations between s^2p^1 and s^2p^6.

The elements in the d-block have valence configurations in which d-orbitals are filled, the elements in the f-block have valence configurations in which f-orbitals are filled, and the elements in the s-block have valence configurations between s^1 and s^2.

4. A

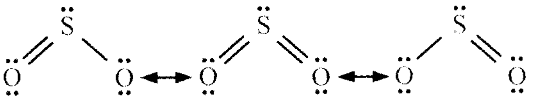
Intermolecular forces affect the melting and boiling points of molecular compounds. There are three types of intermolecular forces:

1. London forces
2. dipole-dipole forces
3. hydrogen bonds

Hydrogen bonds are the strongest intermolecular forces (followed by dipole-dipole, and then London), but they are present only in molecules with O–H, N–H, or H–F bonds. The given table summarizes the attractive forces affecting H_2O, H_2S, H_2Se, and H_2Te. Water has hydrogen bonds in addition to the other two types of intermolecular forces. It therefore takes more energy to vapourize water, which means that water has a higher boiling point than the other substances.

Substance	London Forces	Dipole-Dipole Interactions	Hydrogen Bonds
$H_2O_{(l)}$	✓	✓	✓
$H_2S_{(g)}$	✓	✓	✗
$H_2Se_{(g)}$	✓	✗	✗
$H_2Te_{(g)}$	✓	✗	✗

5. Open Response



Sulfur dioxide has three simultaneous structures because of electron resonance, which means that a pair of electrons moves freely between the bonds, changing the placement of the double bond. At all times, there is one lone pair of electrons on the sulfur atom.

6. B

The atomic number of strontium is 38. Using Hund's rule, the electronic configuration of strontium is represented as $1s^22s^22p^63s^23p^63d^{10}4s^24p^65s^2$.

Krypton is the noble gas that precedes strontium in the periodic table; its atomic number is 36. The shorthand notation of the electronic configuration of strontium is therefore $[\text{Kr}]5s^2$.

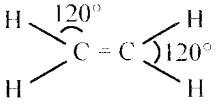
7. Open Response

$\text{Tl}:[\text{Xe}]4f^{14}5d^{10}6s^26p^1$

$\text{Tl}^+:[\text{Xe}]4f^{14}5d^{10}6s^2$

$\text{Tl}^{3+}:[\text{Xe}]4f^{14}5d^{10}$

8. Part A – Open Response

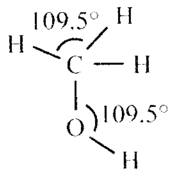


C_2H_4

Ethene

The shape around each carbon atom is trigonal planar.

Part B – Open Response



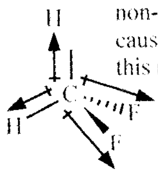
CH_3OH

Methanol

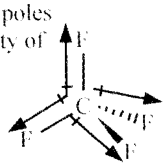
The shape around the carbon atom is tetrahedral, and the shape around the oxygen atom is V-shaped.

9. A

Only molecules that contain polar bonds can be polar, and then only if the bond dipoles do not cancel each other out as a result of molecular geometry. Thus, to predict a molecule's polarity, the shape of a molecule must be determined using the VSEPR theory.

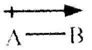


C—F bonds are essentially non-polar. The C—F dipoles cause the overall polarity of this molecule .



The electronegativities of C and S are the same. Hence, no dipoles exist.

The electronegativities of P and H are the same. Hence, no dipoles exist.



or

$^{\delta+}A—B^{\delta-}$

Communicates which atom has the greater share of the bonding electrons.

10. 1 3 5 6

When the geometry of a molecule is asymmetrical, the molecule can have a nonzero dipole moment, making it a polar molecule. Molecules 2 and 4 are symmetrical through all axes of rotation, and therefore are not polar molecules.

11. B

All molecular and atomic chemical entities are subject to London forces, H_2S included. H_2S is a polar covalent compound, so its atoms are also subject to dipole-dipole forces.

The covalent bond also indicates that the valence electrons of the atoms are being shared. In an ionic bond, electrons are lost and gained by constituent atoms, such that they become positive and negative ions that are attracted to one another by virtue of their charge. H_2S does not have any ionic bonds in its structure.

H_2S does not contain hydrogen bonds because at least one of the three most electronegative elements—nitrogen, oxygen, and fluorine—must be present for hydrogen bonds to form.

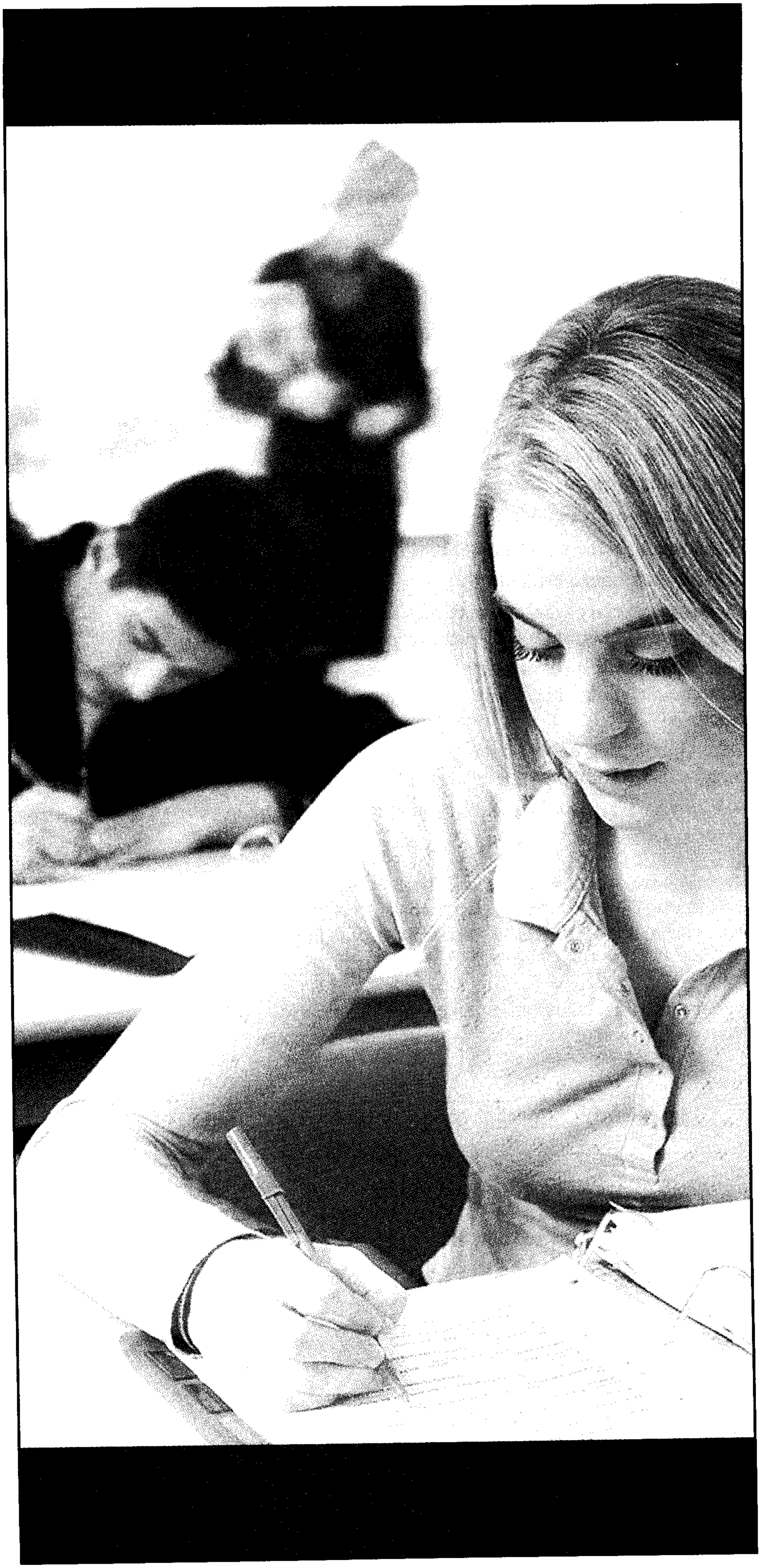
12. Open Response

Compound	Solubility in Water	Solubility in Isopropanol
NaCl	Y	Y
$C_{(graphite)}$	N	N
Wax Shavings	N	Y (slightly)

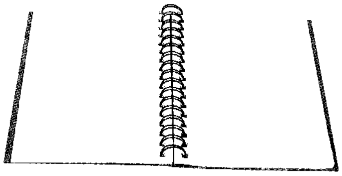
NaCl is an ionic compound. It will dissolve in both solvents because it will be attracted to the polar ends of the molecules in each solvent.

Graphite is a covalent-network compound with strong non-polar intermolecular forces. Its attraction for itself will be greater than its attraction to the non-polar end of isopropanol. The lack of polarity also explains why it will not dissolve in water, which is a polar solvent.

Wax is a non-polar molecular compound. It will not be attracted to water but will be slightly attracted to the non-polar end of isopropanol at standard temperature and pressure. When heated, and the wax shavings become liquefied, the solubility will increase. In solid form, the non-polar intermolecular bonds in wax are stronger than the attractive forces of the isopropanol solvent.



Success on Tests



KEY STRATEGIES FOR SUCCESS ON TESTS

Things to Consider When Taking a Test

- It is normal to feel anxious before you write a test. You can manage this anxiety by:
    - thinking positive thoughts. Visual imagery is a helpful technique to try.
    - making a conscious effort to relax by taking several slow, controlled, deep breaths. Concentrate on the air going in and out of your body.
- Before you begin the test, ask questions if you are unsure of anything.
- Jot down key words or phrases from any oral directions.
- Look over the entire test to assess the number and kinds of questions on the test.
- Read each question closely and reread if necessary.
- Pay close attention to key vocabulary words. Sometimes these are bolded or italicized, and they are usually important words in the question.
- Mark your answers on your answer sheet carefully. If you wish to change an answer, erase the mark completely and then ensure your final answer is darker than the one you have erased.
- On the test booklet, use highlighting to note directions, key words, and vocabulary that you find confusing or that are important to answering the question.
- Double-check to make sure you have answered everything before handing in your test.

When taking tests, the easy words are often overlooked. Failure to pay close attention to these words can result in an incorrect answer. One way to avoid this is to be aware of these words and to underline, circle, or highlight these words while you are taking the test.

Even though some words are easy to understand, they can change the meaning of the entire question, so it is important that you pay attention to them. Here are some examples.

all	**always**	**most likely**	**probably**	**best**	**not**
difference	**usually**	**except**	**most**	**unlikely**	**likely**

Example

1. Which of the following items is **not** considered abiotic?

A. Wind

B. Bacteria

C. Sunlight

D. Precipitation

Helpful Strategies for Answering Multiple-Choice Questions

A multiple-choice question provides some information for you to consider and then asks you to select a response from four choices. Each question has one correct answer. The other answers are distractors, which are incorrect.

Below are some strategies to help you when answering multiple-choice questions.

- Quickly skim through the entire test. Find out how many questions there are and plan your time accordingly.
- Read and reread questions carefully. Underline key words and try to think of an answer before looking at the choices.
- If there is a graphic, look at the graphic, read the question, and go back to the graphic. Then, you may want to underline the important information from the question.
- Carefully read the choices. Read the question first and then each answer that goes with it.
- When choosing an answer, try to eliminate those choices that are clearly wrong or do not make sense.
- Some questions may ask you to select the best answer. These questions will always include words like **best, most appropriate**, or **most likely**. All of the answers will be correct to some degree, but one of the choices will be better than the others in some way. Carefully read all four choices before choosing the answer you think is the best.
- If you do not know the answer or if the question does not make sense to you, it is better to guess than to leave it blank.
- Do not spend too much time on any one question. Make a mark (*) beside a difficult question and come back to it. If you are leaving a question to come back to later, make sure you also leave the space on the answer sheet.
- Remember to go back to the difficult questions at the end of the test; sometimes clues are given throughout the test that will provide you with answers.
- Note any negative words like **no** or **not** and be sure your choice fits the question.
- Before changing an answer, *be sure* you have a very good reason to do so.
- Do not look for patterns on your answer sheet.

About Science Tests

What You Need to Know about Science Tests

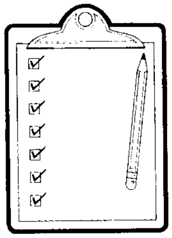
To do well on a science test, you need to understand and apply your knowledge of scientific concepts. Reading skills can also make a difference in how well you perform. Reading skills can help you follow instructions and find key words, as well as read graphs, diagrams, and tables. Math skills are also important to success on science exams. Formulas are often used to solve scientific problems. Practicing with formulas is the best way to become good at them. They can also help you solve science problems.

Science tests usually have two types of questions: questions that ask for understanding of scientific concepts and questions that test how well you can solve scientific problems.

How You Can Prepare for the Science Test

Below are some strategies that are particular to preparing for and writing science tests.

- Know how to use your calculator and, if it is allowed, use your own for the test.
- Note-taking is a good way to review and study important information from your class notes and textbook.
- Sketch a picture of the problem, procedure, or term. Drawing is helpful for learning and remembering concepts.
- Check your answer to practice questions the require formulas by working backward to the beginning. You can find the beginning by going step-by-step in reverse order.
- When answering questions with graphics (pictures, diagrams, tables, or graphs), read the test question carefully.
    - Read the title of the graphic and any key words.
    - Read the test question carefully to figure out what information you need to find in the graphic.
    - Go back to the graphic to find the information you need.
- Always pay close attention when pressing the keys on your calculator. Repeat the procedure a second time to be sure you pressed the correct keys.



TEST PREPARATION COUNTDOWN

There is little doubt that if you develop a plan for studying and test preparation, you *will* perform well on tests.

Below is a general plan to follow seven days before you write a test.

Countdown: 7 Days before the Test

1. Use "Finding Out About the Test" to help you make your own personal test preparation plan.
2. Review the following information:
   - areas to be included on the test
   - types of test items
   - general and specific test tips
3. Start preparing for the test at least 7 days before the test. Develop your test preparation plan and set time aside to prepare and study.

Countdown: 6, 5, 4, 3, 2 Days before the Test

1. Review old homework assignments, quizzes, and tests.
2. Rework problems on quizzes and tests to make sure you still know how to solve them.
3. Correct any errors made on quizzes and tests.
4. Review key concepts, processes, formulas, and vocabulary.
5. Create practice test questions for yourself and then answer them. Work out many sample problems.

Countdown: The Night before the Test

6. The night before the test is for final preparation, which includes reviewing and gathering material needed for the test before going to bed.
7. Most important is getting a good night's rest and knowing you have done everything possible to do well on the test.

Test Day

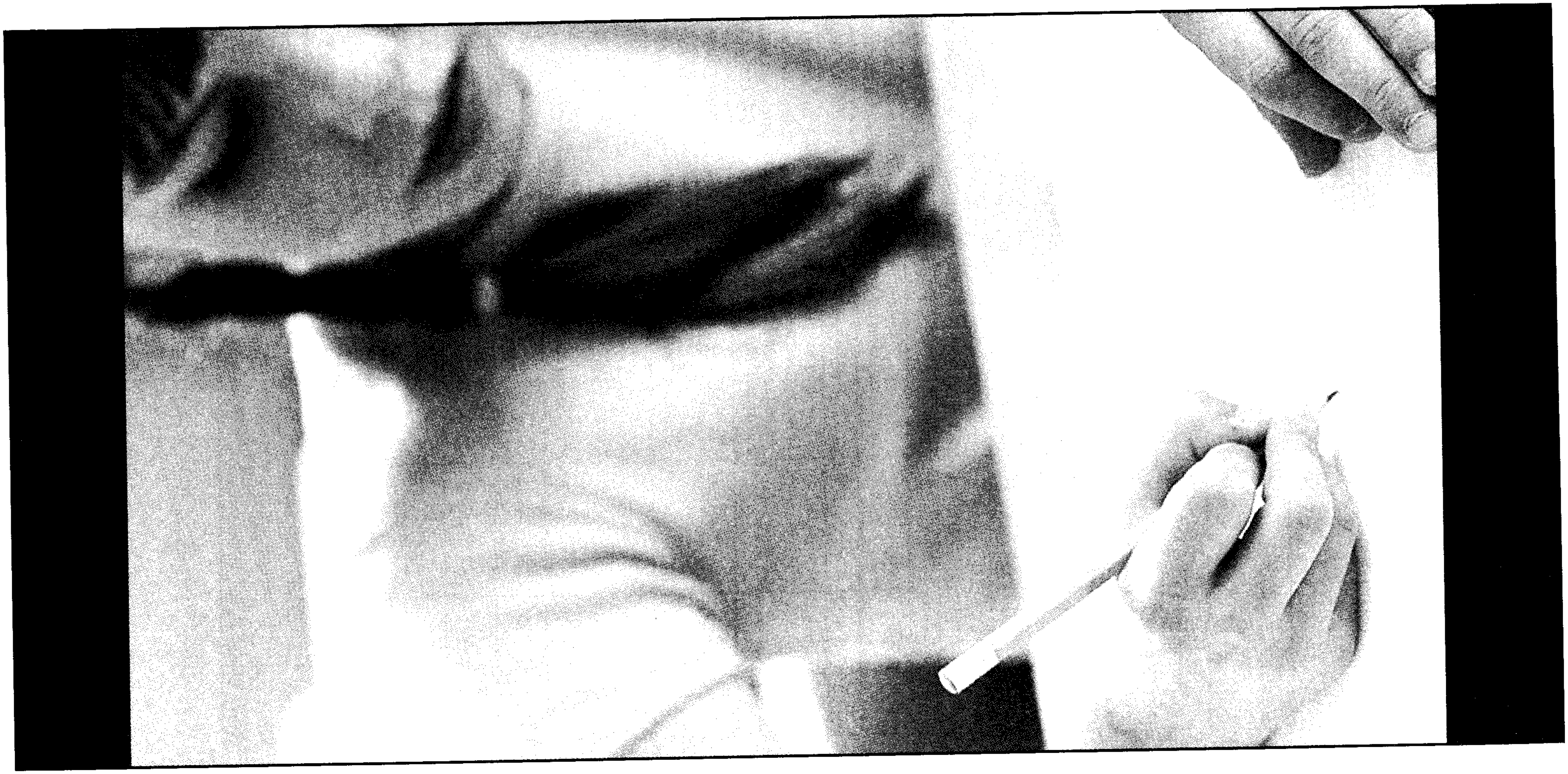
8. Eat a healthy and nutritious breakfast.
9. Ensure you have all the necessary materials.
10. Think positive thoughts: "I can do this." "I am ready." "I know I can do well."
11. Arrive at your school early so you are not rushing, which can cause you anxiety and stress.

SUMMARY OF HOW TO BE SUCCESSFUL DURING THE TEST

The following are some strategies you may find useful for writing your test.

- Take two or three deep breaths to help you relax.
- Read the directions carefully and underline, circle, or highlight any important words.
- Survey the entire test to understand what you will need to do.
- Budget your time.
- Begin with an easy question or a question you know you can answer correctly rather than following the numerical question order of the test.
- If you cannot remember how to answer a question, try repeating the deep breathing and physical relaxation activities first. Then, move to visualization and positive self-talk to get you going.
- Write down anything you remember about the subject on the reverse side of your test paper. This activity sometimes helps you to remind yourself that you *do* know something and you *are* capable of writing the test.
- Look over your test when you have finished and double-check your answers to be sure you did not forget anything.

Practice Tests



Practice Tests

Table of Correlations

Specific Expectation	**Practice Test 1**	**Practice Test 2**
12.1.1 Understanding Basic Concepts		
12.1.1.1 *distinguish among the different classes of organic compounds, including alcohols, aldehydes, ketones, carboxylic acids, esters, ether, amines, and amides, by name and by structural formula*		1
12.1.1.2 *describe some physical properties of the classes of organic compounds in terms of solubility in different solvents, molecular polarity, odour, and melting and boiling points*	1	2
12.1.1.3 *describe different types of organic reactions, such as substitution, addition, elimination, oxidation, esterification, and hydrolysis*	2	
12.1.1.4 *demonstrate an understanding of the processes of addition and condensation polymerization*	3	3
12.1.1.5 *describe a variety of organic compounds present in living organisms, and explain their importance to those organisms*	4	4
12.1.3 Relating Science to Technology, Society, and the Environment		
12.2.1 Understanding Basic Concepts		
12.2.1.1 *compare the energy changes resulting from physical changes, chemical reactions, and nuclear reactions (fission and fusion)*		7
12.2.1.2 *explain Hess's law, using examples*	7	
12.2.1.3 *describe, with the aid of a graph, the rate of reaction as a function of the change of concentration of a reactant or product with respect to time; express the rate of reaction as a rate law equation (first- or second-order reactions only); and explain the concept of half-life for a reaction*		8a, 8b, 8c
12.2.1.4 *analyse simple potential energy diagrams of chemical reactions*	8	
12.2.1.5 *demonstrate understanding that most reactions occur as a series of elementary steps in a reaction mechanism*	9	
12.2.2 Developing Skills of Inquiry and Communication		
12.2.2.2 *write thermochemical equations, expressing the energy change as a ΔH value or as a heat term in the equation*	10a, 10b	9
12.2.2.3 *determine heat of reaction using a calorimeter, and use the data obtained to calculate the enthalpy change for a reaction*	11	10
12.2.2.4 *apply Hess's law to solve problems, including problems that involve data obtained through experimentation*	12	11a, 11b, 11c
12.2.2.5 *calculate heat of reaction using tabulated enthalpies of formation*	13	12
12.2.2.6 *determine through experimentation a rate of reaction, and measure the effect on it of temperature, concentration, and catalysis*	14	13
12.2.3 Relating Science to Technology, Society, and the Environment		
12.2.3.1 *compare conventional and alternative sources of energy with respect to efficiency and environmental impact*		14
12.3.1 Understanding Basic Concepts		
12.3.1.2 *demonstrate an understanding of the law of chemical equilibrium as it applies to the concentrations of the reactants and products at equilibrium*	15	15
12.3.1.3 *demonstrate an understanding of how Le Châtelier's principle can predict the direction in which a system at equilibrium will shift when volume, pressure, concentration, or temperature is changed*	16	16
12.3.1.5 *describe the tendency of reactions to achieve minimum energy and maximum entropy*	17	17

Specific Expectation	Practice Test 1	Practice Test 2
12.3.1.6 *describe, using the concept of equilibrium, the behaviour of ionic solutes in solutions that are unsaturated, saturated, and supersaturated*	18	18
12.3.1.7 *define constant expressions, such as* K_{sp}, K_w, K_a, *and* K_b		19
12.3.1.8 *compare strong and weak acids and bases using the concept of equilibrium*	19	
12.3.1.9 *describe the characteristics and components of a buffer solution*		20
12.3.2 Developing Skills of Inquiry and Communication		
12.3.2.2 *apply Le Châtelier's principle to predict how various factors affect a chemical system at equilibrium, and confirm their predictions through experimentation*	20	21
12.3.2.3 *carry out experiments to determine equilibrium constants*	21	22
12.3.2.4 *calculate the molar solubility of a pure substance in water or in a solution of a common ion, given the solubility product constant* (K_{sp}), *and vice versa*	22, 23	23, 24a, 24b
12.3.2.5 *predict the formation of precipitates by using the solubility product constant*	24a, 24b	
12.3.2.6 *solve equilibrium problems involving concentrations of reactants and products and the following quantities:* K_{eq}, K_{sp}, K_a, K_b, *pH, pOH*	25	25
12.3.2.7 *predict, in qualitative terms, whether a solution of a specific salt will be acidic, basic, or neutral*	26	
12.3.2.8 *solve problems involving acid-base titration data and the pH at the equivalence point*	27	26
12.3.3 Relating Science to Technology, Society, and the Environment		
12.4.1 Understanding Basic Concepts		
12.4.1.1 *demonstrate an understanding of oxidation and reduction in terms of the loss and the gain of electrons or change in oxidation number*		27
12.4.1.2 *identify and describe the functioning of the components in galvanic and electrolytic cells*	28	
12.4.1.3 *describe electrochemical cells in terms of oxidation and reduction half-cells whose voltages can be used to determine overall cell potential*		28
12.4.1.4 *describe the function of the hydrogen half-cell as a reference in assigning reduction potential values*	29	
12.4.1.5 *demonstrate an understanding of the interrelationship of time, current, and the amount of substance produced or consumed in an electrolytic process (Faraday's law)*		29
12.4.2 Developing Skills of Inquiry and Communication		
12.4.2.2 *demonstrate oxidation-reduction reactions through experiments, and analyse these reactions*	30	30
12.4.2.3 *write balanced chemical equations for oxidation-reduction systems, including half-cell reactions*	31	31
12.4.2.4 *determine oxidation and reduction half-cell reactions, direction of current flow, electrode polarity, cell potential, and ion movement in typical galvanic and electrolytic cells, including those assembled in the laboratory*	32	32, 33
12.4.2.5 *predict the spontaneity of redox reactions and overall cell potentials by studying a table of half-cell reduction potentials*	33	
12.4.2.6 *solve problems based on Faraday's law*	34	
12.4.2.7 *measure through experimentation the mass of metal deposited by electroplating, and apply Faraday's law to relate the mass of metal deposited to the amount of charge passed*	35	34

Specific Expectation	Practice Test 1	Practice Test 2
12.4.3 Relating Science to Technology, Society, and the Environment		
12.5.1 Understanding Basic Concepts		
12.5.1.1 *explain the experimental observations and inferences made by Rutherford and Bohr in developing the planetary model of the hydrogen atom*		35
12.5.1.2 *describe the quantum mechanical model of the atom and the contributions of individuals to this model*	36	
12.5.1.3 *list characteristics of the s, p, d, and f blocks of elements, and explain the relationship between position of elements in the periodic table, their properties, and their electron configurations*	37	36
12.5.1.4 *explain how the properties of a solid or liquid depend on the nature of the particles present and the types of forces between them*	38	37
12.5.1.5 *explain how the Valence Shell Electron Pair Repulsion (VSEPR) model can be used to predict molecular shape*	39	38
12.5.2 Developing Skills of Inquiry and Communication		
12.5.2.2 *write electron configurations for elements in the periodic table, using the Pauli exclusion principle and Hund's rule*	40	39
12.5.2.3 *predict molecular shape for simple molecules and ions, using the VSEPR model*	41	40
12.5.2.4 *predict the polarity of various substances, using molecular shape and the electronegativity values of the elements of the substances*	42, 43	41
12.5.2.5 *predict the type of solid (ionic, molecular, covalent network, or metallic) formed by a substance, and describe its properties*		42
12.5.2.6 *conduct experiments to observe and analyse the physical properties of different substances, and to determine the type of bonding present*	44	
12.5.3 Relating Science to Technology, Society, and the Environment		
12.1.2 Developing Skills of Inquiry and Communication		
12.1.2.2 *use the IUPAC system to name and write appropriate structures for the different classes of organic compounds, including alcohols, aldehydes, ketones, carboxylic acids, esters, ethers, amines, amides, and simple aromatic compounds*	5	
12.1.2.3 *build molecular models of a variety of aliphatic, cyclic, and aromatic organic compounds*		5
12.1.2.5 *predict and correctly name the products of organic reactions, including substitution, addition, elimination, esterification, hydrolysis, oxidation, and polymerization reactions*	6	6

Practice Test 1

Practice Test 1

Open Response

1. The boiling point of a molecular compound is a measure of the type and extent of bonding between molecules of that compound.

In an experiment, four unknown organic samples were heated to their boiling points. The data from the experiment is as follows:

Sample	Boiling Point (°C)	Identity
1	17	
2	−38	
3	78.3	
4	−89	

Each sample contains one of the following substances: CH_3CH_3, CH_3CH_2F, $CH_3CH_2NH_2$, CH_3CH_2OH

Using the experimental data, determine the identity of each sample and support your answer with appropriate reasoning.

____

____

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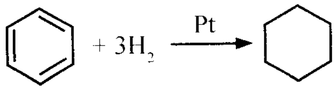
2.



The given reaction is an example of

F. an addition reaction

G. a substitution reaction

H. a combustion reaction

J. an elimination reaction

3. Petroleum is used for a variety of products in addition to the more commonly known uses of gasoline and natural gas. Refineries extract products, such as ethane, from crude oil. Ethane can be chemically converted to ethylene. Ethylene can then be made into polyethylene, a very common plastic used to make materials such as wire insulation and bottles.

What are the two chemical processes that are used to make polyethylene from ethane?

A. Elimination and condensation polymerization

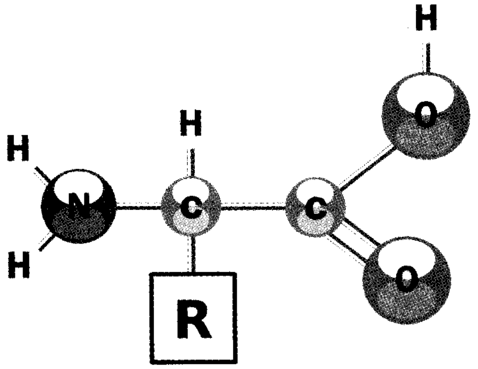
B. Substitution and addition polymerization

C. Elimination and addition polymerization

D. Combustion and substitution

Open Response

4. Amino acids are the building blocks of proteins and can also serve other important functions in the cell. The 20 naturally occurring amino acids are distinguished from one another by their differing functional groups.
The structure of a basic amino acid is given, with the "R" representing the fact that any one of 20 functional groups can be at that place in the structure.



Explain why these molecules are called amino acids.

Open Response

5. Draw a complete structural diagram for 2,2-dimethylbutane.

6. The hydrolysis of CH_3COCl yields
   F. $HCHO + CH_3Cl$

   G. $CH_3CHO + HCl$

   H. $CH_3COOH + HCl$

   J. $HCOOH + CH_3Cl$

7. Two moles of A react with one mole of B to give one mole of C and two moles of D. The energy released during this reaction is 200 kJ. One mole of B reacts with one mole of C to give two moles of E. The energy released during this reaction is 150 kJ.

   How much energy is released when two moles of A react with two moles of B to give two moles each of D and E?
   A. 350 kJ **B.** 300 kJ

   C. 200 kJ **D.** 150 kJ

Open Response

8. Most arenas and curling rinks have artificial ice. Many ice-making plants use ammonia as the refrigerant. The ammonia is circulated in pipes under the ice of the arena or curling rink. For this question, assume that the **only** changes to the ammonia are represented in the equilibrium shown below.

$$NH_{3(l)} \rightleftharpoons NH_{3(g)} \qquad \Delta H = +23.3\text{ kJ}$$

Draw and label an enthalpy diagram that represents the energy changes that occur to the ammonia below the ice surface and to the water on the ice surface as the refrigeration system operates.

9. A reaction has the following mechanism:

Step 1	$2NO \rightarrow N_2O_2$
Step 2	$N_2O_2 + H_2 \rightarrow N_2O + H_2O$
Step 3	$N_2O + H_2 \rightarrow N_2 + H_2O$

Which of the following substances is a reaction intermediate?

A. H_2 **B.** NO

C. H_2O **D.** N_2O

10. Water at 60° C has a $K_w = 9.5 \times 10^{-14}$.

Part A

Open Response

Write an equation representing the ionization of water. Include the heat of reaction (57.1 kJ) in the equation.

Part B

Open Response

If a small amount of NaOH is added to the water, what happens to the value of K_w?

Numerical Response

11. A student heated a 120.0 g sample of $H_2O_{(l)}$ from 21.0°C to 32.5°C by adding 5.93 kJ of energy. The student then used this data to calculate the specific heat capacity of water and compared it with the standard value.
The experimental percentage difference was ____%.
(Record your **three-digit** answer.)

Numerical Response

12. Calculate ΔH for the reaction $CH_{4(g)} + NH_{3(g)} \rightarrow HCN_{(g)} + 3H_{2(g)}$.

Enter your **four-digit** answer here:
____ kJ.

Data:

$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$
$\Delta H = -91.8$ kJ

$C_{(s,\ graphite)} + 2H_{2(g)} \rightarrow CH_{4(g)}$
$\Delta H = -74.9$ kJ

$H_{2(g)} + 2C_{(s,\ graphite)} + N_{2(g)} \rightarrow 2HCN_{(g)}$
$\Delta H = 270.3$ kJ

Open Response

13. A space shuttle uses more than one type of rocket fuel. The two solid rocket boosters use a fuel mixture of aluminium and ammonium perchlorate that reacts according to the following equation:

$$3Al_{(s)} + 3NH_4ClO_{4(s)} \rightarrow Al_2O_{3(s)} + AlCl_{3(s)} + 3NO_{(g)} + 6H_2O_{(g)} + \Delta H$$

In the three main shuttle engines, a mixture of hydrogen and oxygen form a second fuel. The hydrogen and oxygen are carried as compressed liquids in a large tank adjoining the shuttle, and they react to produce energy according to the following reaction:

$$H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(g)}$$

Relevant Heats of Formation

H_f° of $NH_4ClO_{4(s)} = -295.3$ kJ/mol
H_f° of $AlCl_{3(s)} = -705.6$ kJ/mol

Compare the two rocket fuels as energy sources for powering the space shuttle. Your response should also include the following information:

- The calculated energy released for each fuel
- An analysis of the energy-to-mass ratio for each fuel
- Environmental concerns related to each fuel

14. Chemical kinetics is a branch of chemistry that deals with reaction rates, the factors upon which reaction rates depend, and the mechanisms by which reactions take place.

Reaction rates of solids do **not** depend on

A. pressure

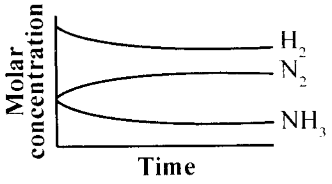
B. temperature

C. the presence of a catalyst

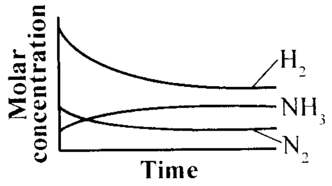
D. the surface area of the reactants

15. Which of the following graphs correctly represents the relationship between elapsed time and the molar concentrations of H_2, N_2, and NH_3 in achieving equilibrium for the reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3$?

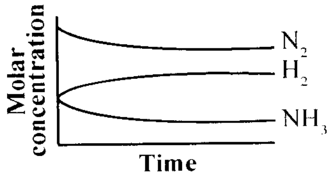
F.



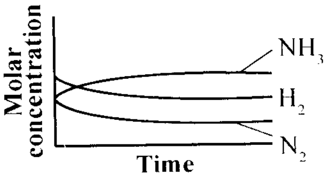
G.



H.



J.



16. Nitric oxide is a catalyst in ozone reduction in the atmosphere, as it reacts with ozone to form nitrogen dioxide. Nitric oxide also causes smog in the lower atmosphere. Nitric oxide is formed in the atmosphere by an endothermic reaction of nitrogen with oxygen, as shown below.

$N_{2(g)} + O_{2(g)} \rightleftharpoons 2NO_{(g)}$

If the temperature of the atmosphere is increased, then the effect on the state of equilibrium of the reaction is that the

A. overall equilibrium of the system will remain unchanged

B. reaction will go to completion, turning all available reactants to products

C. amount of nitric oxide will increase and the amounts of oxygen and nitrogen will decrease

D. amount of nitric oxide will decrease and the amounts of oxygen and nitrogen will increase

17. $4PH_{3(g)} \rightleftharpoons P_{4(g)} + 6H_{2(g)}$ + Heat

The information listed in which of the following charts holds **true** for the given equation?

F.

Enthalpy	**Entropy**	**Inference**
Favours reactants	Favours reactants	Reaction does not occur

G.

Enthalpy	**Entropy**	**Inference**
Favours products	Favours products	Reaction reaches completion

H.

Enthalpy	**Entropy**	**Inference**
Favours reactants	Favours products	Reaction reaches equilibrium

J.

Enthalpy	**Entropy**	**Inference**
Favours products	Favours reactants	Reaction reaches equilibrium

18. A supersaturated solution contains

A. less solute than it would if the dissolved solute were in equilibrium with the undissolved solute

B. more solute than it would if the dissolved solute were in equilibrium with the undissolved solute

C. more solvent than a saturated solution at the same temperature

D. less solvent than a saturated solution at the same temperature

19. The dissociation constants of CH_3NH_2, $(CH_3)_2NH$, NH_3, and NH_2NH_2 are 3.70×10^{-4}, 5.40×10^{-4}, 1.77×10^{-5}, and 1.70×10^{-6}, respectively.

The dissociation constant of the **strongest** base is

F. 1.70×10^{-6}

G. 1.77×10^{-5}

H. 5.40×10^{-4}

J. 3.70×10^{-4}

Open Response

20. Identify a stress that would shift the following equilibrium system to favour the products.

$$2NO_{2(g)} \rightleftharpoons N_2O_{4(g)} + 101.4\text{ kJ}$$

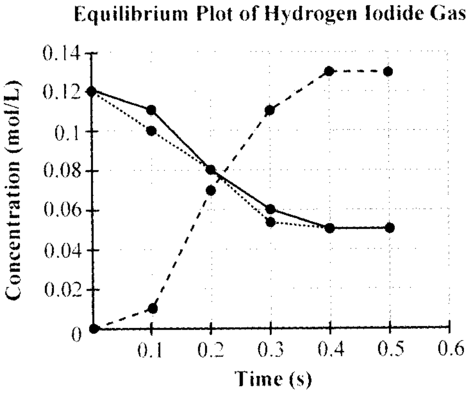
Brown Colourless

Your response should include
- an explanation, based on Le Châtelier's principle, of how the stress would shift this equilibrium
- a procedure that would demonstrate the effect of applying the stress
- a prediction of what evidence there would be that the shift had occurred

Go On

Numerical Response

21. When hydrogen and iodine gas are combined in a closed system, they combine to make hydrogen iodide gas. This experiment was completed at 20°C and the following data was collected. Unfortunately, the graph was not properly labelled, so it is unknown which curve represents which chemical species.



From the graph, determine the K_{eq} of the reverse reaction.

22. The solubility product constant of silver chromate (Ag_2CrO_4) is 1.12×10^{-12}.

Which combination describes the solubility of Ag_2CrO_4 in water compared to a solution of 0.1 M Na_2CrO_4 is correct?

A.

Water	Na_2CrO_4
Increases	Increases

B.

Water	Na_2CrO_4
Remains the same	Decreases

C.

Water	Na_2CrO_4
Decreases	Decreases

D.

Water	Na_2CrO_4
Decreases	Remains the same

Numerical Response

23. The solubility product constant of Ag_2CrO_4 at a very high temperature is 9.0×10^{-9}.

$Ag_2CrO_{4(s)} \rightleftharpoons 2Ag^{+}_{(aq)} + CrO^{2-}_{4(aq)}$

Calculate the concentration of Ag_2CrO_4 in a heated 0.020 M solution of $CrO^{-2}_{4(aq)}$ ___ M. Express your answer as a multiple of 10^{-4}.

24. Sufficient $Na_2(SO_4)_{(s)}$ is added to 0.10 M $Ba(NO_3)_2$ to cause a precipitate to form at standard atmospheric temperature and pressure.

Part A

Open Response

Write the net ionic equation for the precipitate formation.

Part B

Open Response

Calculate the $\left[SO_4^{2-}\right]$ at the moment the precipitate starts to form if the $K_{sp} = 1.1 \times 10^{-10}$ at 25°C.

Numerical Response

25. At a temperature of 300°C and a pressure of 40.5 MPa, 90.0 mol of $H_{2(g)}$ and 80.0 mol of $N_{2(g)}$ are injected into a reaction vessel.
When equilibrium is established, 37.0 mol of $NH_{3(g)}$ are present.

The number of moles of $H_{2(g)}$ present in this equilibrium mixture is ____ mol.

26. When solid potassium cyanide is added to water, the pH of the water

F. increases

G. decreases

H. stays constant

J. becomes neutral

Numerical Response

27. Citric acid, which is a triprotic acid, is completely neutralized by aqueous sodium hydroxide.

If 24.2 mL of 0.100 mol/L $NaOH_{(aq)}$ are required to completely neutralize a 25.0 mL aqueous solution of citric acid, then the acid concentration is ____ mmol/L.
(Record your answer to three digits.)

28. If a galvanic cell consists of solid nickel in a $Ni(NO_3)_2$ solution and solid silver in an $AgNO_3$ solution as the electrodes, which of the following phrases **best** describes what happens in the cell?

F. Nickel is the cathode and nickel ions are oxidized

G. Silver is the cathode and silver ions are reduced

H. Nickel is the anode and nickel ions are reduced

J. Silver is the anode and silver ions are oxidized

Open Response

29. Write the net ionic equation for the redox reaction that takes place in a standard hydrogen-silver cell.

30. Which of the following statements **best** describes why iron(II) nitrate does not react with solid silver?

F. Iron(II) nitrate does not dissociate in solution.

G. Solid silver does not react with anything.

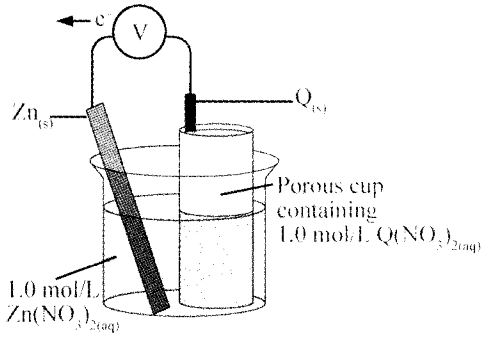
H. The reaction is not spontaneous.

J. Silver nitrate is insoluble.

Open Response

31. Write the balanced ionic equation for the reduction of chlorine gas by aluminium.

32.



Given that the reading on the voltmeter for this cell is +1.74 V, which of the following statements is **true**?

F. $Zn_{(s)}$ is a weaker reducing agent than $Q_{(s)}$.

G. The reduction potential of $Q^{2+}_{(aq)}$ is +2.50 V.

H. $Q^{2+}_{(aq)}$ would react spontaneously with $Cu_{(s)}$.

J. $Q^{2+}_{(aq)}$ is a stronger oxidizing agent than $Zn^{2+}_{(aq)}$.

33. Which of the following reactions will be spontaneous?

A. A silver strip dipped in copper(II) nitrate

B. A gold strip dipped in copper(II) nitrate

C. A silver strip dipped in gold(III) nitrate

D. A gold strip dipped in silver nitrate

Open Response

34. What current would need to be applied to get 1.00 g of gold from a solution of $AuCl_3$ in 1.00 min? Show all of your work.

Numerical Response

35. How many grams of nickel can be plated onto an object if a current of 30.0 A is applied for 2 000.0 s? ____ (Give an answer to three significant digits.)

Open Response

36. Several scientists have contributed to people's understanding of the atom over the last century. What was French physicist Louis de Broglie's major contribution to the current quantum model of the atom? Be sure to include the de Broglie equation in your answer.

$$\lambda = \frac{h}{m \cdot v}$$

37. To which of the following blocks do inert gases belong?

A. d-block

B. f-block

C. p-block

D. s-block

38. The reason metals act as good conductors of electricity is that they have

F. ionic bonding

G. hydrogen bonding

H. filled spaces in the valence energy band

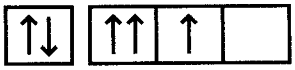
J. empty spaces in the valence energy band

Open Response

39. Formaldehyde, CH_2O, has a trigonal planar shape, yet the bond angle between the two hydrogen atoms is smaller than 120° (the typical bond angle for trigonal planar molecules). Explain this observation.

40. Which of the following orbital diagrams violates **both** the Pauli exclusion principle and Hund's rule of maximum multiplicity?

F.



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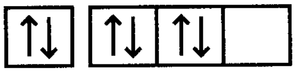

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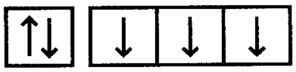
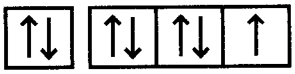
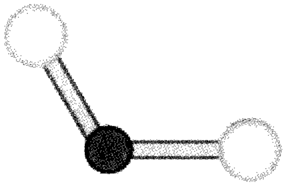


J.

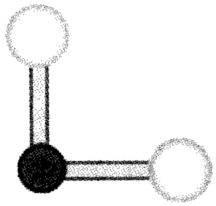


41. Which of the following molecular models represents the molecular structure of CO_2?

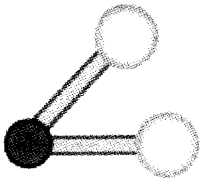
A.



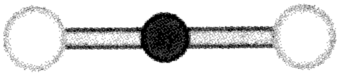
B.



C.



D.



42. Which of the following molecules is a polar molecule?

F. CF_4 **G.** BF_3 **H.** OF_2 **J.** F_2

Numerical Response

43. 1. BeF_2
2. HF
3. NH_3
4. BF_3
5. CCl_4
6. C_2H_2
7. C_2H_4
8. H_2S
9. H_2

In numerical order, the three compounds that are soluble in polar solvents are ____, ____, and ____. (Record your answer.)

Open Response

44. Many solids can be classified and identified by observing their physical properties.

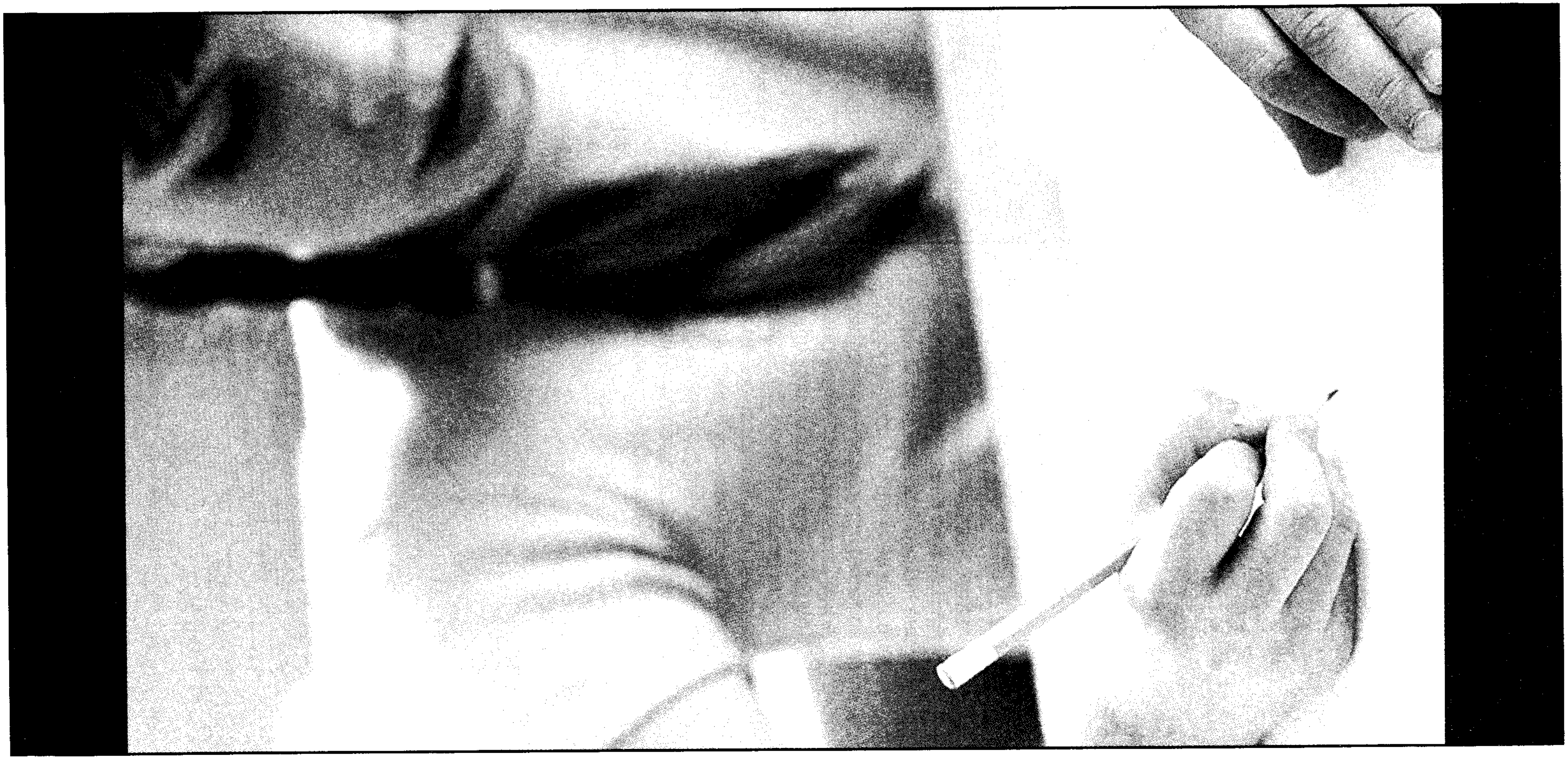
In an experiment, the following data was collected for $CaCO_3$, S, SiC, and Fe. Use the data to identify which set of observations belong to which solid and classify the type of solid.

EC - Electrical Conductivity
MP - Melting Point
M - Malleability
TS - Type of Solid
S - Solid

EC	yes	no	yes (molten)	no
MP	high	low	high	n/a*
M	Malleable	Brittle	Brittle	Hard
TS				
S				

*N/A Not Attainable in class

Practice Test 2



Practice Test 2

1. To which of the following compound classes does R-$CONH_2$ belong?

A. Imides **B.** Amines

C. Amides **D.** Anhydrides

2. Acetaldehyde (CH_3CHO), ethanol (C_2H_5OH), methoxy methane (CH_3OCH_3), and propane (C_3H_8) have comparable molecular masses but varying boiling points. Their molecular masses are 44 g mol^{-1}, 46 g mol^{-1}, 46 g mol^{-1}, and 44 g mol^{-1}, respectively.

If the given organic compounds are arranged in **decreasing** order of their boiling points, the correct order is

F. $CH_3OCH_3 > CH_3CHO > C_2H_5OH > C_3H_8$

G. $C_2H_5OH > CH_3CHO > CH_3OCH_3 > C_3H_8$

H. $C_3H_8 > CH_3OCH_3 > CH_3CHO > C_2H_5OH$

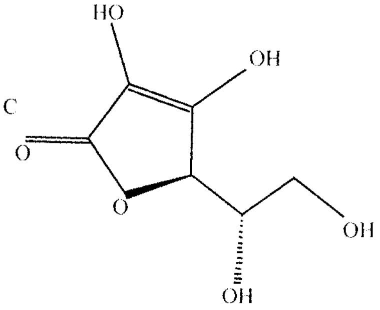
J. $C_2H_5OH > C_3H_8 > CH_3OCH_3 > CH_3CHO$

Open Response

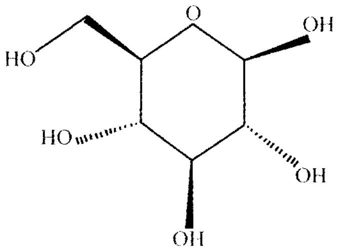
3. Write an equation for the formation of polyethylene from ethylene (ethene). Classify the reaction type.

4.

Vitamin C



Glucose



Vitamin C and glucose are two very important organic molecules in human health. Vitamin C deficiency can cause scurvy, while improper blood glucose levels are associated with diabetes. In order for many biological molecules to function properly, they must be blood soluble, which then enables them to be transferred to nearly every part of the body. This is an important factor in synthetic drug design, and is usually accomplished by giving the drugs certain characteristic structural features.

Which of the following features enables vitamin C and glucose to be blood soluble?

F. Hydroxyl groups

G. Cyclic hydrocarbons

H. Small size of the molecules

J. Oxygen as part of the ring structures

Open Response

5. Name and draw the line structural diagram representing $CH_3CH_2CH_2COOH$.

Open Response

6. Propene can undergo an addition polymerization reaction. Name the product formed from this reaction.

7. 1. $H_2O_{(g)} \rightarrow H_2O_{(l)}$

2. ${}^{3}_{2}He + {}^{3}_{2}He \rightarrow {}^{4}_{2}He + 2p^+$

3. $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$

4. $H_2O_{(l)} \rightarrow H_2O_{(g)}$

Which of the following lists puts the given reactions in order from greatest to lowest **absolute** energy change per mole of reactants?

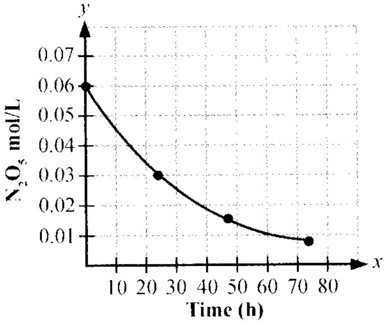
A. 1, 2, 3, 4

B. 4, 1, 2, 3

C. 2, 3; 1 and 4 are the same

D. 3, 2; 1 and 4 are the same

8. N_2O_5 is an unstable gas and decomposes naturally over time. The following graph shows the decomposition of $N_2O_{5(g)}$.



Part A

Open Response

What is the definition of the term *half-life*?

Part B

Open Response

What is the half-life of N_2O_5?

Part C

Open Response

What concentration of N_2O_5 would be expected after 96 hours?

9. If the coefficients of an equation are doubled in a thermochemical equation, then the value of ΔH will be

A. quadrupled **B.** unchanged

C. doubled **D.** halved

Numerical Response

10. A 1.98 g sample of benzoic acid (C_6H_5COOH) is burned in a bomb calorimeter containing excess oxygen. The reaction produces enough heat to raise the temperature of the 1.74 L of water and the 650 g calorimeter by 3.67°C. The specific heat capacity of the calorimeter is 0.924 J / °C.

Determine the molar heat of reaction of benzoic acid, expressed in 10^3 kJ / mol ____.

11. The heat of combustion of magnesium is difficult to determine in a high school chemistry lab. The reaction simply releases too much energy to be measured in a foam cup calorimeter. However, the ΔH_{comb} of magnesium can be derived indirectly using the following reactions:

1. $H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(l)}$
   $\Delta H = -285.5$ kJ / mol
2. $Mg_{(s)} + 2HCl_{(aq)}$
   $\rightarrow MgCl_{2(aq)} + H_{2(g)} \Delta H = ?$
3. $MgO_{(s)} + 2HCl_{(aq)}$
   $\rightarrow H_2O_{(l)} + MgCl_{2(aq)}$
   $\Delta H = -134.9$ kJ / mol
   (Experimentally derived)

An experiment was carried out in which reactions 2 and 3 were performed in foam cup calorimeters. The following data table gives the data obtained from the experiment.

	Experiment 2 with Mg	**Experiment 3 with MgO**
Volume of HCl (mL)	100	100
Mass of colloid (g)	0.5	1.0
T_i(°C)	20	17
T_f(°C)	39	25

Part A

Open Response

Using the experimental data, find the molar enthalpy change for the second reaction. **Hint:** $HCl_{(aq)}$ is dilute and therefore can be assumed to have similar properties to water.

Part B

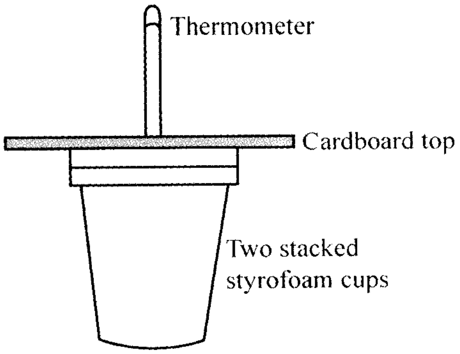
Open Response

Find the molar heat of combustion of magnesium (a.k.a. the heat of formation of magnesium oxide) using Hess's law and the three-step reactions given.

Part C

Open Response

Would you expect the experimental value for the combustion of magnesium to be too high or too low? Explain, given the following diagram for the experimental setup.

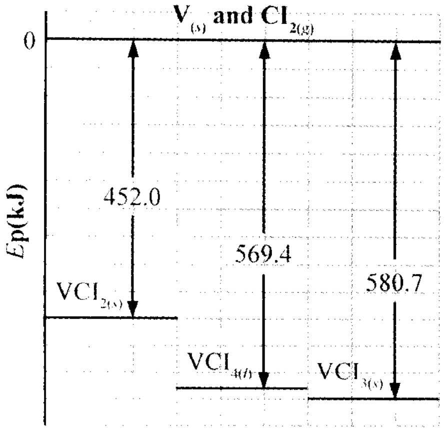


Calorimeter diagram

Numerical Response

12. The following diagram illustrates the formation enthalpies of $V_{(s)}$, $Cl_{2(g)}$, and a selection of their compounds.

Formation Enthalpies of Vanadium Chlorides



The amount of energy absorbed when 0.350 mol of $VCl_{4(l)}$ decomposes to form $VCl_{2(s)}$ and $Cl_{2(g)}$ is ____ kJ.

(Record your **three-digit** answer.)

13. Which of the following statements regarding catalysts is **false**?

A. Catalysts increase the rate of a reaction.

B. Catalysts decrease the activation energy of a reaction.

C. Catalysts increase the number of successful collisions.

D. Catalysts affect the energies of reactants and products.

Open Response

14. Nuclear power plants produce huge amounts of electrical energy — thousands of megawatts per station. Unfortunately, they are not very efficient. The following table gives the efficiency of three types of power stations.

Type of Station	Efficiency (%)
Hydroelectric	70
Fossil fuels	36
Nuclear	30

In a unified and organized essay, discuss the types of energy conversions in each given type of power station. Also, discuss where energy is lost in each type of power station.

15. Iron oxidizes in the presence of water according to the given equation.

$3Fe_{(s)} + 4H_2O_{(g)} \rightleftharpoons Fe_3O_{4(s)} + 4H_{2(g)}$

What is the equilibrium constant of the given reaction in terms of partial pressures?

A. $\dfrac{[Fe_3O_4][H_2]^4}{[Fe]^3[H_2O]^4}$

B. $\dfrac{[H_2O]^4[Fe]^3}{[Fe_3O_4][H_2]^4}$

C. $\dfrac{p_{H_2}^4}{p_{H_2O}^4}$

D. $\dfrac{p_{H_2O}^4}{p_{H_2}^4}$

16. $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$;
$\Delta H = 270.2$ kJ / mol

Which of the following factors will **not** affect the equilibrium of the system?

F. Increase in temperature

G. Decrease in pressure

H. Decrease in the $[I_2]$

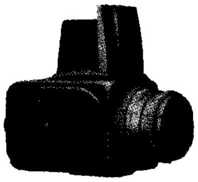
J. Increase in the [HI]

Open Response

17. $Ba(OH)_2 \cdot 8H_2O + NH_4SCN_{(s)} \rightarrow Ba(SCN)_{2(aq)} + H_2O_{(l)} + NH_{3(g)}$
ΔH = positive

Determine if the given unbalanced reaction will be spontaneous. Support your reasoning.

18.



Modern photography uses silver bromide to form images on paper. Silver bromide is slightly soluble in water, and when saturated, the solution is described by the equation $AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)}$.

In this equilibrium equation, the rate of dissolution is

F. equal to the rate of crystallization

G. less than the rate of crystallization

H. greater than the rate of crystallization

J. independent of the rate of crystallization

19. The product of the molar concentrations of H_3O^+ and OH^- ions is known as the

A. ionic product of water

B. solubility product of water

C. molar concentration of hydronium ions

D. dissociation constant of hydronium ions

20. A buffer solution is made by mixing a

F. strong acid with a strong base

G. conjugate salt with a weak base

H. weak acid with a strong base

J. strong acid with a conjugate salt

21. The reaction of carbon monoxide and water takes place as follows:

$CO_{(g)} + H_2O_{(g)} \rightleftharpoons CO_{2(g)} + H_{2(g)}$

Which of the following statements correctly describes the impact of adding $AlCl_3$ as a catalyst on the given state of equilibrium?

A. The amount of CO at equilibrium will decrease.

B. There will be no change in the state of equilibrium.

C. The equilibrium will shift toward the forward direction.

D. The equilibrium will shift toward the backward direction.

Numerical Response

22. The acid dissociation constant, K_a, for acetic acid can be found experimentally using a 0.100 M solution of CH_3COOH and a pH metre. When the pH of the solution is no longer changing, this information is used to determine the K_a of acetic acid. This dissociation proceeds as follows:

$CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$

If the pH of the acetic acid solution was found to be 2.850, what is the K_a of the acid? (record your three-digit answer) ____ $\times 10^{-5}$.

Numerical Response

23. The pH of a saturated aqueous solution of $Ba(OH)_2$ is 10.0. If the K_{sp} of $Ba(OH)_2$ is 1×10^{-9}, then the concentration of Ba^{2+} ions in the solution in mol / L is ____Ba^{2+}.

24. Beryllium phosphate is formed by the reaction of phosphoric acid with a beryllium salt. The reaction is exothermic and produces hydrogen gas.

Part A

Open Response

Write the chemical equilibrium equation to illustrate the equilibrium that exists in a saturated solution of beryllium phosphate.

Part B

Open Response

Give the K_{sp} expression for this compound.

Numerical Response

25. Sodium azide, which is found in automobile airbags, reacts readily with acids to form the highly toxic and explosive hydroazoic acid $HN_{3(aq)}$. The K_a for hydroazoic acid is 1.9×10^{-5}.

The pH of a 0.28 mol/L $HN_{3(aq)}$ solution is ____. (Record your **three-digit** answer.)

Open Response

26. What is the pH of the equivalence point after titrating 100 mL of 1.00 M NaClO with 2.00 M HCl? The K_b for NaClO is 3.6×10^{-7}.

27. $2Fe^{2+}_{(aq)} + H_2O_{2(aq)} + 2H^{+}_{(aq)} \rightarrow 2Fe^{3+}_{(aq)} + 2H_2O_{(l)}$

How many electrons are transferred in the given redox reaction?

F. 0 **G.** 1 **H.** 2 **J.** 3

Numerical Response

28. The half-reaction $Na^{+}_{(aq)} + e^{-} \rightarrow Na_{(s)}$ has a reduction potential of −2.37 V, and the half-reaction $Cl_{2(g)} + 2e^{-} \rightarrow 2Cl^{-}_{(aq)}$ has a reduction potential of 1.36 V.

What voltage must be applied to allow the following cell to function at standard conditions:

$C_{(s)} \mid Cl_{2(g)} \mid Cl^{-}_{(aq)} \parallel$

$Na^{+}_{(aq)} \mid Na_{(s)} \mid C_{(s)}$? ____ V

Open Response

29. What current is necessary to reduce 1.00 mol of Au^{3+} to $Au_{(s)}$ in 1.00 h?

Open Response

30. An experimental setup contains an unknown metal, a solution of hydrochloric acid, and a solution of nickel(II) nitrate.

Explain, using the table of standard reduction potentials as a reference, which experimental results would **best** identify the unknown metal.

Go On

31. The number of electrons present in the balanced half-reaction for the reduction of bromine gas to bromide is

F. 1 electron on the product side

G. 1 electron on the reactant side

H. 2 electrons on the product side

J. 2 electrons on the reactant side

32. In an electrolytic cell, two inert carbon electrodes are placed in an acid solution of sodium chloride. What half-reaction occurs at the anode?

A. $Na^{+}_{(aq)} + e^{-} \rightarrow Na_{(s)}$

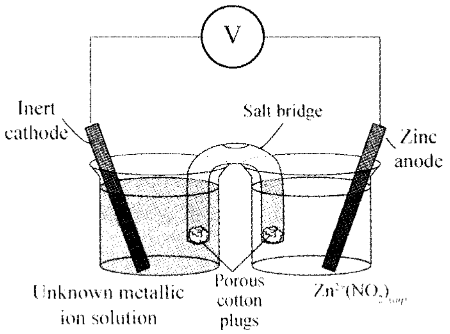
B. $H_{2(g)} \rightarrow 2H^{+}_{(aq)} + 2e^{-}$

C. $2Cl^{-}_{(aq)} \rightarrow Cl_{2(g)} + 2e^{-}$

D. $2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^{+}_{(aq)} + 4e^{-}$

Open Response

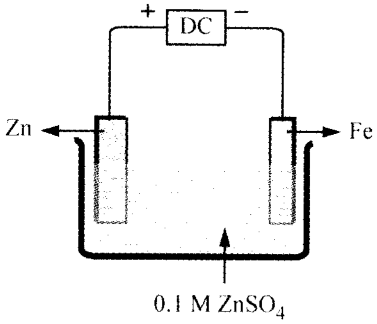
33.



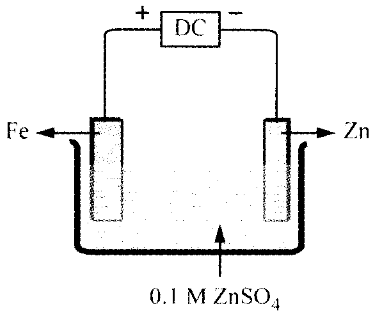
Explain why the unknown solution cannot contain $Cr(NO_3)_2$ in a functioning galvanic cell.

34. Which of the following electrolytic cells can be used for electroplating iron with zinc metal?

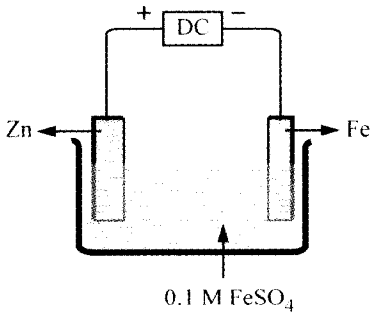
A.



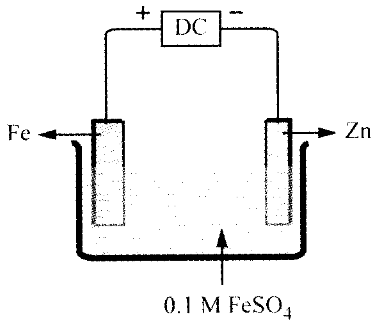
B.



C.



D.



35. Which of the following statements explains a weakness in the planetary model of the atom?

F. The negative electron would be attracted by the positive nucleus with a Coulomb force sufficient to collapse the atom.

G. The centripetal force acting on the electron exceeds the electrostatic force of attraction between electron and nucleus.

H. The calculated orbital velocity could not be counterbalanced by the centripetal force due to the electrostatic attraction between the electron and nucleus.

J. An orbiting electron would be in a state of constant acceleration, so it should radiate energy continuously, reducing its kinetic energy until it spirals into the nucleus.

Open Response

36. In an alternate universe, the following periodic table was discovered with a unique set of elements specific to that universe:

Alternate Periodic Table

s block					
1	2			p block	
3	4			5	6
7	8	d block		9	10
11	12	13	14	15	16
17	18	19	20	21	22

In this universe, the quantum number m_l has been changed, but the remaining quantum values are the same.

$n = 1, 2, 3, 4\ldots$

$l = 0, 1, 2\ldots(n - 1)$

$m_l = ?$

$m_s = +\frac{1}{2}, -\frac{1}{2}$

What is the new value of m_l in this universe? Justify your answer.

37. Hydrogen cyanide, a highly poisonous compound, undergoes addition reactions with organic compounds, such as the aldehydes and ketones, to form a compound known as cyanohydrin.

The intermolecular bonding in HCN involves

F. ionic bonding and dispersion forces

G. hydrogen bonding and dispersion forces

H. dispersion forces and dipole-dipole forces

J. hydrogen bonding and dipole-dipole forces

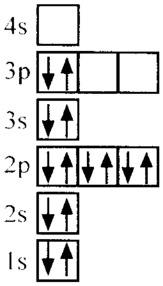
38. Among the molecules that have the same shape, the bond angle around the central atom changes with the change in its electronegativity. Which of the following isostructural compounds has the **smallest** bond angle?

A. H_2O **B.** H_2S

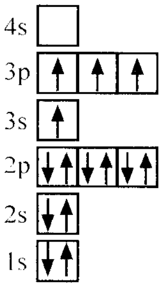
C. H_2Se **D.** H_2Te

39. Which orbital box diagram represents silicon?

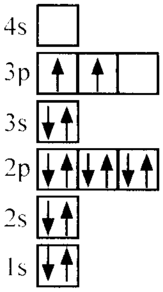
F.



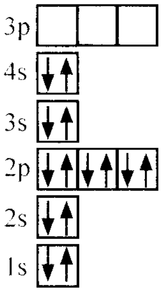
G.



H.



J.



Open Response

40. Fill in the following table by determining the molecular shape for each molecule. Use the VSEPR method.

Molecule	3-D VSEPR Diagram	Name of Shape
PCl_3		
HCN		
SF_6		
CO_3^{2-}		

Numerical Response

41. 1. H_2S

2. PF_3

3. TeF_6

4. SiF_4

Of the given four molecules, which two would be non-polar, based on molecular geometry? ____, ____.

42.  Chlorine (Cl_2) is used in great quantity as a bleach for the pulp and paper industry and as a disinfectant for municipal water supplies. Iodine is another chemical commonly used for water disinfection on a smaller scale. The boiling point of chlorine is −34.6°C, while the boiling point of iodine is 184°C.

The difference in boiling points between chlorine and iodine can be attributed to

A. ionic bonds

B. covalent bonds

C. intramolecular bonds

D. London dispersion forces

SOLUTIONS–PRACTICE TEST 1

1. OR
2. F
3. C
4. OR
5. OR
6. H
7. A
8. OR
9. D
10. Part A- **OR**
Part B- **OR**
11. 2.56
12. 256.0
13. OR
14. A
15. G
16. C
17. G
18. B
19. H
20. OR
21. 0.15
22. B
23. 3.4
24. Part A- **OR**
Part B- **OR**
25. 34.5
26. F
27. 32.3
28. G
29. OR
30. H
31. OR
32. F
33. C
34. OR
35. 18.3
36. OR
37. C
38. J
39. OR
40. F
41. D
42. H
43. 2,3,8
44. OR

1. Open Response

Sample	Boiling Point (°C)	Identity
1	17	$CH_3CH_2NH_2$
2	−38	CH_3CH_2F
3	78.3	CH_3CH_2OH
4	−89	CH_3CH_3

Ethane, CH_3CH_3, has the fewest intermolecular bonds (only London dispersion) and so will have the lowest boiling point. Fluoroethane has both London dispersion and dipole-dipole intermolecular forces, making it the second lowest boiling point. The boiling point of ethanol exceeds that of ethanamine because hydroxyl functional groups tend to be stronger hydrogen bonded than N-H functional groups: the -OH bond is more polar. Both ethanol and ethanamine will have boiling points higher than that of fluoroethane.

2. F

The given reaction is an addition reaction because nothing is removed and the unsaturated benzene ring is converted into saturated cyclohexane.

3. C

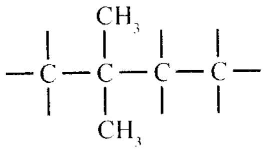
Ethane can undergo a process called cracking, which is an elimination reaction. Then polyethylene is made from the addition polymerization of ethylene monomers.

4. Open Response

The name amino acid comes from the two functional groups that make up the backbone of the molecule. In the given figure, the left side of the molecule is an amine group and the right side is a carboxylic acid group. The amine and the acid make up the name amino acid.

5. Open Response

2,2-dimethylbutane
(C_6H_{14})



6. H

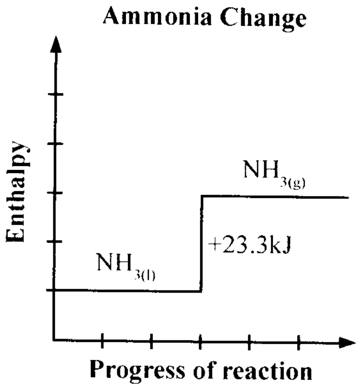
The hydrolysis of CH_3COCl is represented as $CH_3COCl + HOH \rightarrow CH_3COOH + HCl$.

Thus, the products obtained are CH_3COOH and HCl.

7. A

The given reactions can be written as follows:
$2A + B \rightarrow C + 2D + 200$ kJ
$B + C \rightarrow 2E + 150$ kJ
The reaction obtained on adding these two reactions is as follows:
$2A + B + B + C \rightarrow C + 2D + 200 \text{ kJ} + 2E + 150$ kJ
or $2A + 2B \rightarrow 2D + 2E + 350$ kJ
Thus, 350 kJ of energy is released when two moles of A react with two moles of B to give two moles each of D and E.

8. Open Response



9. D

A reaction intermediate is a substance formed in one step in the reaction mechanism that is consumed in a later step.

10. Part A – Open Response

For example,

$2H_2O_{(l)} + 57.1\text{ kJ} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$ OR

$H_2O_{(l)} + 57.1\text{ kJ} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$

Note: The endothermic nature of the reaction can be deduced from the data provided.

Part B – Open Response

For a small change in $[OH^-]$, K_w remains unchanged.

11. 2.56

For a kinetic energy change (only), the energy involved, q, is readily calculated from $q = mc\Delta t$.

In this case, $q = 5.93$ kJ, $m = 0.1200$ kg, and $\Delta t = (32.5 - 21.0)°C = 11.5°C$.

Thus,

$$c = \frac{q}{m\Delta t}$$
$$= \frac{5.93\text{ kJ}}{0.1200\text{ kg} \times 11.5°C}$$
$$= \frac{4.30\text{ kg}}{\text{kg}°C} \text{ or } \frac{4.30\text{ J}}{\text{g}°C}$$

The percentage difference is given by % difference.

$$= \frac{\text{reference value} - \text{measured value}}{\text{reference value}} \times 100\%$$
$$= \frac{\frac{4.19\text{ J}}{\text{g}°C} - \frac{4.30\text{ J}}{\text{g}°C}}{\frac{4.19\text{ J}}{\text{g}°C}} \times 100\%$$
$$= 2.56\%$$

12. 256.0

Using Hess' law,

$2NH_{3(g)} \rightarrow N_{2(g)} + 3H_{2(g)}$

$\Delta H = 91.8$ kJ

$[CH_{4(g)} \rightarrow C_{(s,\ graphite)} + 2H_{2(g)}$

$\Delta H = 74.9\text{ kJ}] \times 2$

$H_{2(g)} + 2C_{(s,\ graphite)} + N_{2(g)} \rightarrow 2HCN_{(g)}$

$\Delta H = 270.3$ kJ

Yield: $[2CH_{4(g)} + 2NH_{3(g)} \rightarrow 2HCN_{(g)} + 6H_{2(g)}$

$\Delta H = 511.9\text{ kJ}] \div 2$

$= CH_{4(g)} + NH_{3(g)} \rightarrow HCN_{(g)} + 3H_{2(g)}$

$\Delta H = 256.0$ kJ

13. Open Response

Energy comparison

$3Al_{(s)} + 3NH_4ClO_{4(s)} \rightarrow Al_2O_{3(s)} + AlCl_{3(s)} + 3NO_{(g)} + 6H_2O_{(g)}$

$\Delta H = [1\text{ mol}(-1\,675.7\text{ kJ/mol}) + 1\text{ mol}(-705.6\text{ kJ/mol}) + 3\text{ mol}(+91.3\text{ kJ/mol}) + 6\text{ mol}(-241.8\text{ kJ/mol})] - [3\text{ mol}(-295.3\text{ kJ/mol})]$

$\Delta H = -2\,672.3$ kJ

(Energy released by the solid rocket fuel is 2 672.3 kJ.)

$$H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(g)} \quad \Delta H = -241.8\text{ kJ}$$

(Energy released by the hydrogen-oxygen fuel is 241.8 kJ.)

Energy released per gram of $Al_{(s)}/NH_4ClO_{4(s)}$ fuel is as follows:

$$\frac{2\,672.3\text{ kJ}}{3\text{ mol}(26.98\text{ g/mol}) + 3\text{ mol}(117.50\text{ g/mol})}$$
$$= 6.17\text{ kJ/g fuel (or } 6.1653286\text{ kJ/g)}$$

Energy released per gram of hydrogen-oxygen fuel is as follows:

$$\frac{241.8\text{ kJ}}{1\text{ mol}\left(2.02\text{ g}/\text{mol}\right) + \frac{1}{2}\text{ mol}\left(32.00\text{ g}/\text{mol}\right)}$$
$$= 13.4\text{ kJ/g fuel (or } 13.42\text{ kJ/g)}$$

Environmental concerns

The solid rocket fuel has the following properties:

- It produces $NO_{(g)}$, which is an atmospheric pollutant and precursor to acid rain that can also deplete the ozone layer.
- It produces solid aluminium compounds, which are atmospheric particulates.
- It uses $Al_{(s)}$, the production of which causes environmental damage.

The hydrogen-oxygen fuel has the following properties:

- It is more environmentally friendly because it produces $H_2O_{(g)}$.
- It requires hydrogen produced directly or indirectly from fossil fuels.

14. A

The rate of a reaction does not depend on pressure.
The rate of a reaction depends on temperature, the surface area of the reactants, and the presence of a catalyst.

15. G

The molar concentrations of H_2 and N_2 decrease proportionally with elapsed time, and at equilibrium, their concentrations become constant. The NH_3 concentration rise is twice the drop in N_2 concentration.

16. C

The given reaction is endothermic in nature. According to Le Châtelier's principle, on increasing the temperature of a system, the equilibrium will move in such a direction that the system loses heat to the surroundings. That is, the equilibrium will shift in the forward direction.

17. G

Heat is evolved in the given reaction. This means that the energy of the products is lesser than that of the reactants. Hence, enthalpy favours the products. In addition to this, four moles of reactants form seven moles of products. This means that there is an increase in entropy. Therefore, entropy also favours the products.
Since both enthalpy and entropy favour the products, the reaction will reach completion.

18. B

At a constant temperature and pressure, the quantity of solute in a saturated solution is also constant.
A supersaturated solution is a solution that contains more of the dissolved solute than could be dissolved under normal circumstances.

19. H

The base with the highest value of the dissociation constant will be the strongest.
Therefore, the dissociation constant of the strongest base is 5.4×10^{-4}.

20. Open Response

Each of the samples that follow represent only one valid approach to each of the problems.
During the diploma examination marking session, provision is made for considering various approaches the students may have used.

Valid Stresses

- add $NO_{2(g)}$
- remove $N_2O_{4(g)}$
- increase pressure or decrease volume
- lower temperature

Sample Procedure

Observe the colour of the gas in a $NO_{2(g)}$–$N_2O_{4(g)}$ equilibrium tube at room temperature. The bulb could then be put in a cold water bath and the colour change observed

Explanation of Le Châtelier's Principle

When the bulb is put in the cold water, it will turn a lighter brown because the removal of energy will make the exothermic reaction predominate, which will make the equilibrium shift to produce more $N_2O_{4(g)}$ (colourless).

Predicted Evidence

- brown colour diminishes
- temperature increases
- pressure decreases

21. 0.15

$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

$$K_{eq} \text{ for the reverse reaction } = \frac{[H_2][I_2]}{[HI]^2}$$

$$= \frac{(0.05)^2}{(0.13)^2} = 0.15$$

Data values are obtained from the graph, and although the curves are not labelled, it should be clear which is which, given that the forward reaction is producing HI. Therefore, the curve increasing is HI, and the other two are the reactants.

22. B

Water does not affect the solubility of Ag_2CrO_4, but Na_2CrO_4, because of the presence of common ions (chromate), depresses the solubility of Ag_2CrO_4. Therefore, the solubility of Ag_2CrO_4 will remain the same in water but will be lower in a Na_2CrO_4 solution.

23. 3.4

$Ag_2CrO_{4(s)}$	$\rightleftharpoons$	$2Ag^+_{(aq)}$	+	$CrO^{2-}_{4(aq)}$
x		$2x$		$0.020 \text{ M} + x$

$K_{sp} = [Ag^+]^2[CrO_4^{2-}]$

$9.0 \times 10^{-9} = (2x)^2(x + 0.020)$

Assume $0.020 + x \cong 0.020$

Check the assumption using the 100 rule:

$\frac{0.020 \text{ M}}{9.0 \times 10^{-9}} > 100$; thus, the assumption is correct.

$$4x^2(0.020) = 9.0 \times 10^{-9}$$
$$0.08x^2 = 9.0 \times 10^{-9}$$
$$x = 3.4 \times 10^{-4} \text{ M}$$

Therefore, because x and Ag_2CrO_4 are in a 1:1 ratio, the concentration of $Ag_2CrO_4 = 3.4 \times 10^{-4}$ M.

24. Part A – Open Response

$Ba^{2+}_{(aq)} + SO^{2-}_{4(aq)} \rightarrow BaSO_{4(s)}$

Part B – Open Response

$$\left[SO_4^{2-}\right] = \frac{K_{sp}}{\left[Ba^{2+}\right]}$$

$$= \frac{1.1 \times 10^{-10}}{0.10}$$

$$= 1.1 \times 10^{-9}$$

25. 34.5

A simple equilibrium calculation is all that is necessary to answer this question.

	$N_{2(g)}$ +	$3H_{2(g)} \rightleftharpoons$	$2NH_{3(g)}$
Initial	80.0 mol	90.0 mol	0 mol
Change	$-x$	$-3x$	$+2x$
Equilibrium	$80.0 - x$	$90.0 - 3x$	$2x$

At equilibrium, $n_{NH_3} = 37.0 \text{ mol} = 2x$

Therefore, $x = 18.5 \text{ mol}$

Consequently,
$$\begin{aligned} n_{H_2} &= 90.0 \text{ mol} - 3x \\ &= 90.0 \text{ mol} - 3(18.5 \text{ mol}) \\ &= 34.5 \text{ mol} \end{aligned}$$

The pressure and temperature information suggest a more complicated question—calculating K_{eq}; however, this was not asked and so this data is irrelevant.

26. F

Water is a neutral compound. Adding solid potassium cyanide (KCN) to water results in the formation of KOH and HCN. KOH is a strong base and HCN is a weak acid, so the solution becomes alkaline.

27. 32.3

The **stoichiometric method** is widely employed to solve for the masses, volumes, and concentrations of reactants or products in chemical reactions. Most teachers would agree that the stoichiometric method comprises the following four steps.

1. Write a balanced reaction equation that includes all the relevant data (often the most difficult step).
2. Calculate the number of moles of the given reactant or product (n_{given}).
3. Use the appropriate mole ratio to convert n_{given} into the number of moles of required reactant or product (n_{required}).
4. Determine the answer by converting the n_{required} into a mass, a volume, or a concentration.

The titration of citric acid, often written as $H_3Ct_{(aq)}$ by $NaOH_{(aq)}$, is represented readily by the following equation:

$$H_3Ct_{(aq)} + 3NaOH_{(aq)} \rightarrow Na_3Ct_{(aq)} + 3HOH_{(l)}$$

$$n_{NaOH} = 24.2 \text{ mL} \times 0.100 \text{ mol/L} / \text{L} = 2.42 \text{ mmol}$$

$$n_{Na_3Ct} = \frac{1}{3} n_{NaOH} = \frac{1}{3} \times 2.42 \text{ mmol} = 0.807 \text{ mmol}$$

$$\left[H_3Ct_{(aq)}\right] = \frac{0.807 \text{ mmol}}{25.0 \text{ mL}} = 0.0323 \text{ mol/L}$$

The citric acid concentration is 0.0323 mol/L. (This can be reported as 32.3 mmol/L.)

28. G

Reduction occurs at the cathode, and oxidation occurs at the anode. The strongest reducing agent (which gets oxidized) acts as the anode and the strongest oxidizing agent (which gets reduced) acts as the cathode. Silver is the stronger oxidizing agent and therefore undergoes reduction at the cathode.

29. Open Response

The standard reduction potential for silver is 0.80 V, which means it would act as the cathode in the hydrogen-silver cell. Therefore, its reaction must be the reduction half-reaction and the hydrogen half-reaction must be the oxidation. The two half reactions are

$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$ and $H_{2(g)} \rightarrow 2H^+_{(aq)} + 2e^-$.

This means that the silver half-reaction must be multiplied by two before they can be combined into the following net ionic equation: $2Ag^+_{(aq)} + H_{2(g)} \rightarrow 2Ag_{(s)} + 2H^+_{(aq)}$.

30. H

Only spontaneous redox reactions proceed without any added energy.

31. Open Response

The two half-reactions are $Cl_2 \rightarrow Cl^-$ and $Al \rightarrow Al^{3+}$.

First, the half-reactions must be balanced:

$Cl_2 + 2e^- \rightarrow 2Cl^-$ and $Al \rightarrow Al^{3+} + 3e^-$

The lowest common multiple is 6, so $3Cl_2 + 6e^- \rightarrow 6Cl^-$ and $2Al \rightarrow 2Al^{3+} + 6e^-$.

Combining the two half-reactions gives

$2Al + 3Cl_2 \rightarrow 6Cl^- + 2Al^{3+}$.

32. F

In the cell diagram, the electron flow makes the zinc electrode act as the cathode of this voltaic cell. In other words, it is the electrode that has a positive $E°_{net}$.

The $Q_{(s)}$ electrode is therefore the anode. The anode half-cell holds the most readily oxidized entity (in other words, the strongest reducing agent in this cell). $Zn_{(s)}$ is therefore a weaker reducing agent than $Q_{(s)}$.

Normally, the cathode is the electrode of a voltaic cell that has the highest reduction potential.

If $Q^{2+}_{(aq)}$ was the stronger oxidizing agent in the cell, the electron flow would enter the $Q_{(s)}$ electrode/cathode.

If the reduction potential of $Q^{2+}_{(aq)}$ was greater than that of copper, it would react spontaneously with copper and zinc, and it would be the cathode of the voltaic cell in the diagram. This is plainly not so.

33. C

Only in the case of silver dipped in a solution of Au^{3+} is the oxidizing agent above the reducing agent on the standard reduction potential table.

34. Open Response

First, find the number of moles of gold: $1.00\ g \times \dfrac{1\ mol}{196.97\ g} = 5.08 \times 10^{-3}\ mol$

The half-reaction for the reduction of gold is $Au^{3+}_{(aq)} + 3e^{-} \rightarrow Au_{(s)}$, so three electrons are needed for each atom of gold. Therefore, the amount of charge needed is $\left(5.08 \times 10^{-3}\ mol\right) \times 3$

$= 1.52 \times 10^{-2}\ mol \times \dfrac{9.647 \times 10^{4}\ C}{mol}$.

To find the current, just divide the charge by time: $\dfrac{1.47 \times 10^{3}\ C}{60.0\ s} = 24.5\ A$

35. 18.3

$30.0\dfrac{C}{s} \times 2\ 000.0\ s = 6.00 \times 10^{4}\ C$

The half-reaction for the reduction of nickel is $Ni^{2+}_{(aq)} + 2e^{-} \rightarrow Ni_{(s)}$.

The moles of electrons are

$6.00 \times 10^{4}\ C \times \dfrac{1\ mol}{9.647 \times 10^{4}\ C} = 0.622\ mol$, which corresponds to 0.311 moles of nickel.

Then convert this to grams: $0.311\ mol \times 58.69\dfrac{g}{mol} = 18.3\ g$

36. Open Response

Louis de Broglie extended Einstein's theory on the wave-particle duality of light to matter. He theorized that particles of matter could show wave properties under certain circumstances. He developed the idea that a particle with mass m and speed v would have an associated wavelength λ by the following equation:

$$\lambda = \frac{h}{m \cdot v}$$

Though the wavelength of most matter is undetectable because it is so small (the mass of most objects is too big and the speed too low), electrons are an exception. They are extremely light and move so quickly that they create detectable waves. Since electrons can act as waves and waves can be described using mathematics, de Broglie had set the stage for quantum mechanics or wave mechanics, which is essentially the use of math to describe the properties and location of an electron within an atom.

37. C

Inert gases belongs to the p-block. Except for helium, the electronic configuration of the outermost orbitals of these elements varies from ns^2np^1 to ns^2np^6. These are called the p-block elements because their properties depend on the presence of the p-electron.

38. J

Metals have empty spaces in their valence energy band. When an electron is excited, it can easily move into the empty spaces and move around freely. These free electrons are responsible for the good electrical and thermal conductivity of metals.

Metals exhibit metallic bonding, not ionic bonding.

Metals exhibit metallic bonding, not hydrogen bonding.

There are no filled spaces in the valence energy band of solids.

39. Open Response

Multiple bonds require more space than single bonds because of the greater number of electrons. The repulsion of double-bonded oxygen is greater than single-bonded hydrogen electrons. Therefore, the hydrogen atoms are forced closer together.

40. F

The orbital diagram represented in alternative F violates both Hund's rule of maximum multiplicity and the Pauli exclusion principle because two electrons occupy the first subshell of the p-orbital prior to the complete filling of the orbital. Moreover, the spins of the two electrons present in the first subshell of the p-orbital are the same, which is not possible.

The orbital diagram represented in alternative G violates only Hund's rule.

The orbital diagrams given in alternatives H and J do not violate any rule.

41. D

Carbon dioxide (CO_2) is tri-atomic molecule made up of two atoms of oxygen and one atom of carbon. It is linear in shape with a bond angle of 180°. There are no pairs of electrons out of the 180° plane.

42. H

Only OF_2 has a non-symmetrical shape, creating an uneven distribution of charge in which the F atoms are slightly negative and the O atom is slightly positive.

43. 2,3,8

Generally speaking, a compound is polar only if it contains polar bonds. However, this is not a sufficient condition for a molecule to be polar. VSEPR theory can predict whether individual bond dipoles give a zero vector sum or not. If the net dipole moment is zero, then the compound is said to be non-polar. Since polar solvents accommodate polar compounds easily, only the following three compounds from the given list are soluble in polar solvents:

HF, NH_3, H_2S

44. Open Response

EC - Electrical Conductivity
MP - Melting Point
M - Malleability
TS - Type of Solid
S - Solid

EC	yes	no	yes (molten)	no
MP	high	low	high	n/a*
M	Malleable	Brittle	Brittle	Hard
TS	Metallic	Molecular	Ionic	Covalent Network
S	Fe	S	$CaCO_3$	SiC
*N/A Not Attainable in class				

SOLUTIONS–PRACTICE TEST 2

1. C	9. C	17. **OR**	26. **OR**	36. **OR**
2. G	10. -1.78	18. F	27. H	37. H
3. **OR**	11. Part A- **OR**	19. A	28. 3.73	38. D
4. F	Part B- **OR**	20. H	29. **OR**	39. H
5. **OR**	Part C- **OR**	21. B	30. **OR**	40. **OR**
6. **OR**	12. 41.1	22. 2.02	31. J	41. 3, 4
7. C	13. D	23. 0.1	32. C	42. D
8. Part A- **OR**	14. **OR**	24. Part A- **OR**	33. **OR**	
Part B- **OR**	15. C	Part B- **OR**	34. A	
Part C- **OR**	16. G	25. 2.64	35. J	

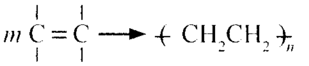
1. C

The compound **R**-$CONH_2$ is the common chemical formula for an amide. Acid amides are compounds in which the hydroxyl of the carboxyl group is replaced by the amino group NH_2 to form the amide group $CONH_2$.

2. G

Ethanol undergoes intermolecular hydrogen bonding. Since it exists as an associated molecule, it has the highest boiling point. Acetaldehyde has a stronger dipole-dipole interaction than methoxy methane. Hence, it will have a higher boiling point. Methoxy methane has a higher dipole-dipole interaction than propane because propane has no hydrogen bonding. Therefore, the correct decreasing order is:
$C_2H_5OH > CH_3CHO > CH_3OCH_3 > C_3H_8$.

3. Open Response



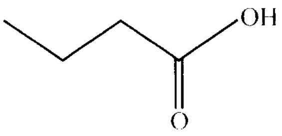
$$m\ \overset{|}{\underset{|}{\mathrm{C}}}=\overset{|}{\underset{|}{\mathrm{C}}} \longrightarrow (\mathrm{CH_2CH_2})_n$$

This is an example of an addition polymerization reaction.

4. F

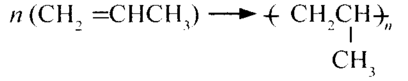
The polar functional groups (in this case the hydroxyl groups) are responsible for the solubility. Blood is made up of mostly water, a polar solvent. Therefore, any organic molecule that functions in an aqueous state must be polar. Functional groups also contribute to interactions with other biomolecules and are largely responsible for the activity of the compound. The other factor for activity is the shape of the molecule, largely contributed by ring and chain structures.

5. Open Response



Butanoic acid

6. Open Response



$$n(\mathrm{CH_2{=}CHCH_3}) \longrightarrow (\mathrm{CH_2}\underset{\underset{\mathrm{CH_3}}{|}}{\mathrm{CH}})_n$$

The product of this reaction is polypropylene.

7. C

Reaction 2 has the greatest energy change because it is a nuclear reaction. Reaction 3 is a chemical combustion, which has a greater energy change than the physical change of state of water (reactions 1 and 4).

8. Part A – Open Response

The time needed for a reactant mass or concentration to decrease by half its initial value.

Part B – Open Response

The half-life of N_2O_5 is 24 h, as shown in the given graph.

Part C – Open Response

96 h ÷ 24 h = 4 half-lives

$$0.06\ \mathrm{mol/L} \times \left(\frac{1}{2}\right)^4 = 0.00375\ \mathrm{mol/L}$$

$$= 3.75 \times 10^{-3}\ \mathrm{mol/L}$$

9. C

If the coefficients of an equation are multiplied by an integer, then ΔH is also multiplied by the same integer. Hence, the value of ΔH will be doubled.

10. -1.78

$m_{\text{water}} = 1.74\text{ L} \times 1.0\text{ g/mL} = 1\,740\text{ mL} \times 1.0\text{ g/mL}$
$= 1\,740\text{ g}$
$q_{\text{rxn}} = -[q_{\text{water}} + q_{\text{cal}}]$
$= -[(m_{\text{water}} \times c_{\text{water}} \times \Delta t_{\text{water}}) + (m_{\text{cal}} \times c_{\text{cal}} \times \Delta t_{\text{cal}})]$
$= [(1\,740\text{ g} \times 4.184\text{ J/g°C} \times 3.67\text{°C})$
$+(650\text{ g} \times 0.924\text{ J/g°C} \times 3.67\text{°C})]$
$= -28.9\text{ kJ}$

$$\Delta H_{\text{benzoic acid}} = \frac{-28.9\text{ kJ}}{1.98\text{ g} \div 122.0\text{ g/mol}}$$
$= -1\,780.7\text{ kJ/mol} = -1.78 \times 10^3\text{ kJ/mol}$

11. Part A – Open Response

Assume $d_{\text{acid}} = d_{\text{water}}$, and assume $c_{\text{water}} = c_{\text{acid}}$.
$m_{\text{acid}} = 100\text{ mL} \times 1\text{g/mL} = 100\text{ g}$
$q_{\text{rxn}} = -(mc\Delta T) = -(100\text{ g} \times 4.184\text{ J/g°C} \times 19\text{°C})$
$= -7\,949.6\text{ J} = -7.95\text{ kJ}$
$\Delta H = -7.95\text{ kJ} \div n_{\text{Mg}} = 7.95\text{ kJ}/(0.5\text{ g} \div 24.3\text{ g/mol})$
$= -386.4\text{ kJ/mol}$

Part B – Open Response

$(1)\text{H}_{2(g)} + \frac{1}{2}\text{O}_{2(g)} \rightarrow \text{H}_2\text{O}_{(l)} \Delta H = -285.5\text{ kJ/mol}$
$(2)\text{Mg}_{(s)} + 2\text{HCl}_{(aq)} \rightarrow \text{MgCl}_2(aq) + \text{H}_{2(g)}$
$\Delta H = -386.4\text{ kJ/mol}$
$(3)\text{H}_2\text{O}_{(l)} + \text{MgCl}_{2(aq)} \rightarrow \text{MgO}_{(s)} + 2\text{HCl}_{(aq)}$
$\Delta H = +134.9\text{ kJ/mol}$
$\text{Mg}_{(s)} + \frac{1}{2}\text{O}_{2(g)} \rightarrow \text{MgO}_{(s)} \Delta H = -537\text{ kJ/mol}$

Part C – Open Response

The combustion of magnesium would be too low.
Heat lost to the air and the calorimeter are not accounted for in this experiment, making final temperatures too low; therefore, ΔT is low.

12. 41.1

It is necessary to write this equation to solve the question. The described equation is $\text{VCl}_{4(l)} \rightarrow \text{VCl}_{2(s)} + \text{Cl}_{2(g)}$
Using H_fs from the graph and the formula
$\Delta H = \sum nH^{\circ}_{f(p)} - \sum nH^{\circ}_{f(r)}$
$\Delta H = (1\text{ mol} \times -452.0\text{ kJ/mol}) -$
$(1\text{mol} \times -569.4\text{ kJ/mol} + 0) = 117.4\text{ kJ}$
$H_{r(\text{VCl}_{4(l)})} = +117.4\text{ kJ/mol}$
$\Delta H = nH = 0.350\text{ mol} \times (+117.4\text{ kJ/mol})$
$= +41.1\text{ kJ}$

13. D

Catalysts increase the rate of a reaction by decreasing the energy of activation for the reaction. However, it does not affect the energy of the reactants or products.
Catalysts decrease the energy of activation by increasing the number of successful or effective collisions.

14. Open Response

In a unified and organized essay, discuss the following points:

- Types of energy conversions in each given type of power station
- Where energy is lost in each type

Your organized essay should include some of the following information.

- Hydroelectric—A hydroelectric power station uses moving water to turn a generator that produces electrical energy. The water has velocity and therefore kinetic energy. This kinetic energy is then used to turn a generator, giving it mechanical kinetic energy.
  The generator then produces electrical energy. Most of the energy lost will be as a result of friction, heat, and sound produced in the generator.
- Fossil fuels—In this type of power station, natural gas or coal is burned in a furnace to turn water into steam. The steam is then used to turn a generator, which produces electrical energy. The natural gas or coal has chemical potential energy that is burned and converted into thermal energy, which is then used to heat water and turn it into steam. The steam is then used to turn a generator giving it mechanical kinetic energy, which then produces electrical energy. Most of the energy is lost up the smokestacks as heat, but some is lost as friction, heat, and sound in the generator.
- Nuclear—In a nuclear reactor, the energy stored in an atom of uranium is released to give off heat that is used to turn water into steam. The steam is then used to turn a generator, creating electricity. The energy in the uranium is released from the nucleus and is, therefore, nuclear energy. It is converted into thermal energy, which is used to turn the water into steam. The steam is then used to turn a generator, giving it mechanical kinetic energy,which produces electrical energy. Energy is also lost as heat exhaust, and in the generator, as friction, heat, and sound energy. However, in a nuclear plant, the conversion from chemical potential to thermal energy is less efficient because of the danger of the nuclear reaction getting out of control, which could result in a meltdown.

15. C

Applying the law of mass action on the given reaction,

r_1 = rate of forward reaction $= K_f[Fe]^3[H_2O]^4$

r_2 = rate of backward reaction $= K_b[Fe_3O_4][H_2]^4$

At equilibrium,

$r_1 = r_2$

Equilibrium constant $= K = \frac{K_f}{K_b} = \frac{[Fe_3O_4][H_2]^4}{[Fe]^3[H_2O]^4}$

Since Fe_3O_4 and Fe are solids, $[Fe_3O_4]$ and [Fe] are taken as constant and equal to 1.

Therefore, $K = \frac{[H_2]^4}{[H_2O]^4}$

Representing the equilibrium constant in terms of partial pressures,

$K_p = \frac{p^4_{H_2}}{p^4_{H_2O}}$, in which $p^4_{H_2}$ and $p^4_{H_2O}$ are partial pressures of H_2 and H_2O, respectively.

16. G

In the given reaction, 2 moles of reactants combine to form 2 moles of products, i.e., there is no change in the number of moles. Hence, according to Le Châtelier's principle, a change in pressure will not affect the equilibrium of the system.

The forward reaction is endothermic. An increase in temperature will favour the forward reaction, which proceeds by absorbing heat.

A decrease in the concentration of I_2 will cause the equilibrium to shift in the reverse direction, which leads to an increase in the concentration of the reactants, according to Le Châtelier's principle.

An increase in the [HI] will shift the equilibrium in the backward direction.

17. Open Response

$Ba(OH)_2 \cdot 8H_2O + 2NH_4SCN_{(s)} \rightarrow$

$Ba(SCN)_{2(aq)} + 10H_2O_{(l)} + 2NH_{3(g)} \Delta H = 170 \text{ kJ}$

$\Delta H > 0$, ∴ not spontaneous

$\Delta S > 0$, ∴ spontaneous

Therefore, according to the Gibbs free energy change equation, $\Delta G = \Delta H - T\Delta S$, the reaction will be spontaneous at high temperatures.

$\Delta G < 0$ at high T.

18. F

In all solutions in which a saturated aqueous solution and undissolved solid solute occupy the same vessel, the following dynamic solubility equilibrium occurs:

$\text{solute}_{(s)} \rightleftharpoons \text{solute}_{(aq)}$

Equilibria like this occur when the rate of the forward process,

$\text{solute}_{(s)} \rightarrow \text{solute}_{(aq)}$(dissolution),

is equal to rate of reverse process,

$\text{solute}_{(s)} \leftarrow \text{solute}_{(aq)}$ (crystallization).

For the solubility equilibrium

$AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)}$,

the solute dissociates when it dissolves, as do all ionic compounds.

19. A

The ionization of water takes place as

$2H_2O \rightleftharpoons H_3O^+ + OH^-$.

According to the law of chemical equilibrium,

$K = \{[OH^-][H_3O^+]\}/[H_2O]$

$\Rightarrow K[H_2O] = [OH^-][H_3O^+]$

$K_w = [OH^-][H_3O^+]$ (∴ The concentration of water is constant.)

K_w is known as the ionic product of water.

20. H

A buffer solution is defined as a solution that resists any change in its pH value even when small amounts of acid or base are added to it. Buffer solutions consist of a weak acid and its conjugate base. It can also contain a weak base and its conjugate acid. Adding a strong base to a solution of weak acid would produce the conjugate base of the weak acid, producing a buffer solution given that the strong base does not entirely neutralize the acid.

21. B

The presence of a catalyst does not affect the equilibrium of a reaction. Catalysts will, however, cause a dynamic reaction to reach equilibrium more quickly.

22. 2.02

$\text{pH} = -\log[\text{H}^+]$

$2.850 = -\log[\text{H}^+]$

$10^{-2.850} = [\text{H}^+] = 0.001412 \text{ M}$

	CH_3COOH	H^+	CH_3COO^-
I	0.100 M	0	0
C	−0.001412	+0.001412	+0.001412
E	0.09859 M	0.001412 M	0.001412 M

$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = \frac{(0.001412)^2}{0.09859} = 2.02 \times 10^{-5}$$

23. 0.1

$\text{Ba(OH)}_2 = \text{Ba}^{2+} + 2\text{OH}^-$

pH = 10 means that $[\text{H}^+] = 10^{-10}$

$[\text{H}^+][\text{OH}^-] = 10^{-14}$ (Ion product of water)

$(10^{-10})[\text{OH}^-] = 10^{-14}$

$[\text{OH}^-] = 10^{-4}$

$K_{sp} = [\text{Ba}^{2+}][\text{OH}^-]^2$ (in which K_{sp} is the solubility product and $[\text{Ba}^{2+}]$ and $[\text{OH}^-]$ are the respective concentrations in the solution)

Therefore,

$10^{-9} = [\text{Ba}^{2+}](10^{-4})^2$

$[\text{Ba}^{2+}] = 10^{-1} \text{ mol/L} = 0.1 \text{ mol/L}$

24. Part A – Open Response

$\text{Be}_3(\text{PO}_4)_{2(s)} \rightleftharpoons 3\text{Be}^{2+}_{(aq)} + 2\text{PO}^{3-}_{4(aq)}$

Part B – Open Response

$K_{sp} = [\text{Be}^{2+}]^3[\text{PO}_4^{3-}]^2$

25. 2.64

For the equilibrium

$\text{HN}_{3(aq)} \rightleftharpoons \text{H}^+_{(aq)} + \text{N}^-_{3(aq)}$

Initially	0.28 mol/L	– –	—
At equilibrium	0.28 − x	x	x

$$K_a = \frac{x^2}{0.28 - x} \cong \frac{x^2}{0.28} \text{ since } \frac{[\text{HN}_3]}{K_a} > 1\,000$$

$x = [\text{H}^+_{(aq)}] = \sqrt{K_a(0.28)}$

$[\text{H}^+_{(aq)}] = \sqrt{1.9 \times 10^{-5} \text{ mol/L} \times 0.28 \text{ mol/L}}$

$= 2.3 \times 10^{-3} \text{ mol/L}$

$\text{pH} = -\log(2.3 \times 10^{-3} \text{ mol/L}) = 2.64$

Whether a quadratic or approximation method is used to compute the $[\text{H}^+_{(aq)}]$, the pH is 2.64.

26. Open Response

$\text{NaClO}_{(aq)} + \text{HCl}_{(aq)} \rightarrow \text{NaCl}_{(aq)} + \text{HClO}_{(aq)}$

After the equivalence point, all that is left in the solution is $\text{HClO}_{(aq)}$, which is a strong conjugate acid. Since the concentration of HCl is twice that of NaClO, only half the volume of HCl would be needed to titrate the sample. The final volume of the solution at the equivalence point is 100 mL NaClO + 50 mL HCl = 150 mL.

Therefore, $C_{\text{HClO}} = \dfrac{1.0 \text{ M} \times 0.1 \text{ L}}{0.15 \text{ L}} = 0.67 \text{ M}.$

$HClO_{(aq)}$	+	$H_2O_{(l)}$	⇌	$H_3O^+_{(aq)}$	+	$ClO^-_{(aq)}$
0.67 M				≅ 0		0
$-x$				$+x$		$+x$
$0.67 - x$				x		x

$K_w = K_a \times K_b$

$$\text{So, } K_a = \frac{1.0 \times 10^{-14}}{3.6 \times 10^{-7}} = 2.8 \times 10^{-8}$$

$$K_a = 2.8 \times 10^{-8} = \frac{[\text{H}_3\text{O}^+][\text{ClO}^-]}{[\text{HClO}]} = \frac{x^2}{0.67 - x}$$

Assume $0.67 - x \cong 0.67$

Check $\dfrac{0.67}{2.8 \times 10^{-8}} > 100$; therefore, the assumption is correct.

$$2.8 \times 10^{-8} = \frac{x^2}{0.67}$$

$x = \sqrt{1.9 \times 10^{-8}} = 1.37 \times 10^{-4} \text{ M} = [\text{H}^+]$

$\text{pH} = -\log[\text{H}^+] = -\log[\text{H}_3\text{O}^+] = 3.87$

27. H

This is a complicated redox equation, but it can be seen that the iron goes from a 2+ state to a 3+ state.
This involves a one-electron oxidation, but in this case, the stoichiometry calls for double that amount, so the answer is two electrons.

28. 3.73

This is an electrolytic cell that involves the reduction of Na^+ and the oxidation of Cl^- to produce $Na_{(s)}$ and $Cl_{2(g)}$. The sum of the reduction potential of one half-reaction and the oxidation potential of the other half-reaction gives the total cell potential required.
Since only the reduction potential of the Cl_2 half-reaction is given, the oxidation potential must be the opposite value of that number or −1.36 V. Thus, the total cell potential is $-2.37 + (-1.36) = -3.73$. This means that 3.73 V must be applied.

29. Open Response

It takes 3.00 mol of electrons to reduce 1.00 mol of Au^{3+}. That is

$$9.647 \times 10^4 \frac{\text{C}}{\text{mol}} \times 3.00\text{ mol} = 2.89 \times 10^5\text{ C and}$$

$$\frac{2.89 \times 10^5\text{ C}}{3\,600\text{ s}} = 80.4\text{ A.}$$

30. Open Response

With the available materials, two main experiments can be conducted. First, some of the unknown metal could be added to the hydrochloric acid. If a reaction occurs, it means that the reduction half-reaction of the unknown metal is below the $2H^+_{(aq)} + 2e^- \rightleftharpoons H_{2(g)}$ half-reaction on the table. If there is no reaction, it is above. Next, the nickel(II) nitrate solution could be applied to a piece of the metal. If a reaction occurs, it means that the unknown metal was able to replace nickel in the nitrate, and therefore Ni^{2+} was reduced. This would mean that the $Ni^{2+}_{(aq)} + 2e^- \rightleftharpoons Ni_{(s)}$ is above the reduction half-reaction of the unknown metal on the table. No reaction means it is below. Since the $2H^+_{(aq)} + 2e^- \rightleftharpoons H_{2(g)}$ and the $Ni^{2+}_{(aq)} + 2e^- \rightleftharpoons Ni_{(s)}$ are near each other on the table with $2H^+_{(aq)} + 2e^- \rightleftharpoons H_{2(g)}$ being higher, the best experimental result would be a reaction with hydrochloric acid and no reaction with nickel(II) nitrate. This would mean the reduction half-reaction of the unknown metal is between the two on the table, which eliminates the most possibilities.

31. J

First, write the half reaction $Br_2 \rightarrow Br^-$. To balance the bromine, write $Br_2 \rightarrow 2Br^-$. This means there is a charge of −2 on the product side, so 2 electrons must be added to the reactant side.

32. C

Oxidation occurs at the anode, so the species most likely to oxidize will react at the anode. The solution contains Na^+, Cl^-, H_2O, and H^+. The products of the electrolysis of brine (saltwater) are hydrogen and chlorine gases, so the half-reaction $2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$ occurs at the anode.

33. Open Response

Since the electrode is inert, it means that only what is available in the metallic solution can undergo the redox reaction. This leaves water, NO_3^-, and Cr^{2+}. Using the standard reduction potential table, it is clear that the solid zinc will undergo the oxidation reaction, as it is lowest on the right. This makes it the anode, and the inert electrode must be the cathode. Reduction occurs at the cathode, so in order for the cell to function, something in the solution must be higher on the left of the table than the zinc ion. Cr^{2+} is not higher.

34. A

The metal that has to be electroplated is made the cathode, and the metal with which plating has to be done is made the anode. Also, a salt of the metal that is used for plating is used as the electrolyte.
Hence, to electroplate iron metal with zinc, iron must be made the cathode and zinc must be made the anode.
Since the negative side of the DC generator is at the iron electrode, it is the cathode. A salt of zinc, zinc sulfate in this case, should be used as the electrolyte. The electrolytic cell given in alternative A can be used for electroplating iron metal with zinc.

35. J

If all the positive matter in the atom resides in a tiny cluster in the atom's centre separated from the negative particles (components of cathode rays), the negative particles can have orbital velocities large enough to counteract the force of attraction between the positive and negative charges.
However, the orbiting charged particles would undergo constant centripetal acceleration. According to Maxwell's theories, an accelerating charge radiates EMR. Therefore, the orbiting charged particle should be emitting energy as it experiences centripetal acceleration. Thus, according to classical electromagnetic theory, the electron would be losing kinetic energy, so its orbital velocity would decrease, making it unable to counteract the electrostatic force. The particle would not be stable and would spiral into the nucleus.

36. Open Response

$m_l = 1$

The value for m_l is unaffected by the value of l as observed in the given periodic table. There is only one of each type of suborbital. There is only one s, one p, and one d orbital, making each block of the periodic table only two columns wide, since only two electrons fit in the s, p, d orbitals in this new universe. Basically, because all of the blocks are two columns wide, there are only two electrons in every orbital level, which corresponds to a level 1 orbital shape. If the blocks were all six columns wide, that would correspond to a level 2 orbital shape.

37. H

For any molecule, the extent and type of intermolecular bonding depends on the nature of the molecules involved. All molecules experience dispersion forces. HCN is a polar molecule, so it is affected by dipole-dipole forces. Only those molecules that contain O-H, N-H, or H-F bonds are affected by hydrogen bonding. Hydrogen bonding in HCN is not possible because the hydrogen atom is bonded to carbon, which has a low electronegativity. Ionic bonds involve electron transfer from one atom to another, whereas HCN, being a covalent compound, involves electron sharing. Thus, HCN experiences only dispersion forces and dipole-dipole forces.

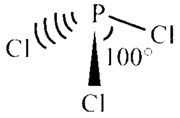
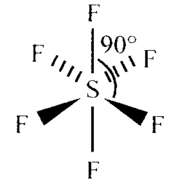
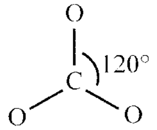
38. D

Among the given compounds, Te (in H_2Te) is the least electronegative. Hence, the electron pairs around Te lie far from it. As a result, the repulsive interaction between the electron pairs decreases. This decreases its bond angle.

39. H

According to Hund's rule of maximum multiplicity and the Pauli exclusion principle, the electronic configuration of silicon is $1s^2 2s^2 2p^6 3s^2 3p^2$.

40. Open Response

Molecule	3-D VSEPR Diagram	Name of Shape
PCl_3		Trigonal pyramidal
HCN	$H-C\equiv N$	Linear
SF_6		Octahedral
CO_3^{2-}		Trigonal planar

41. 3, 4

Molecules 3 and 4 will be non-polar, as they are symmetrical shapes. Tellurium hexafluoride will assume an octahedral form, which is best imagined as two square pyramids joined together at their bases.

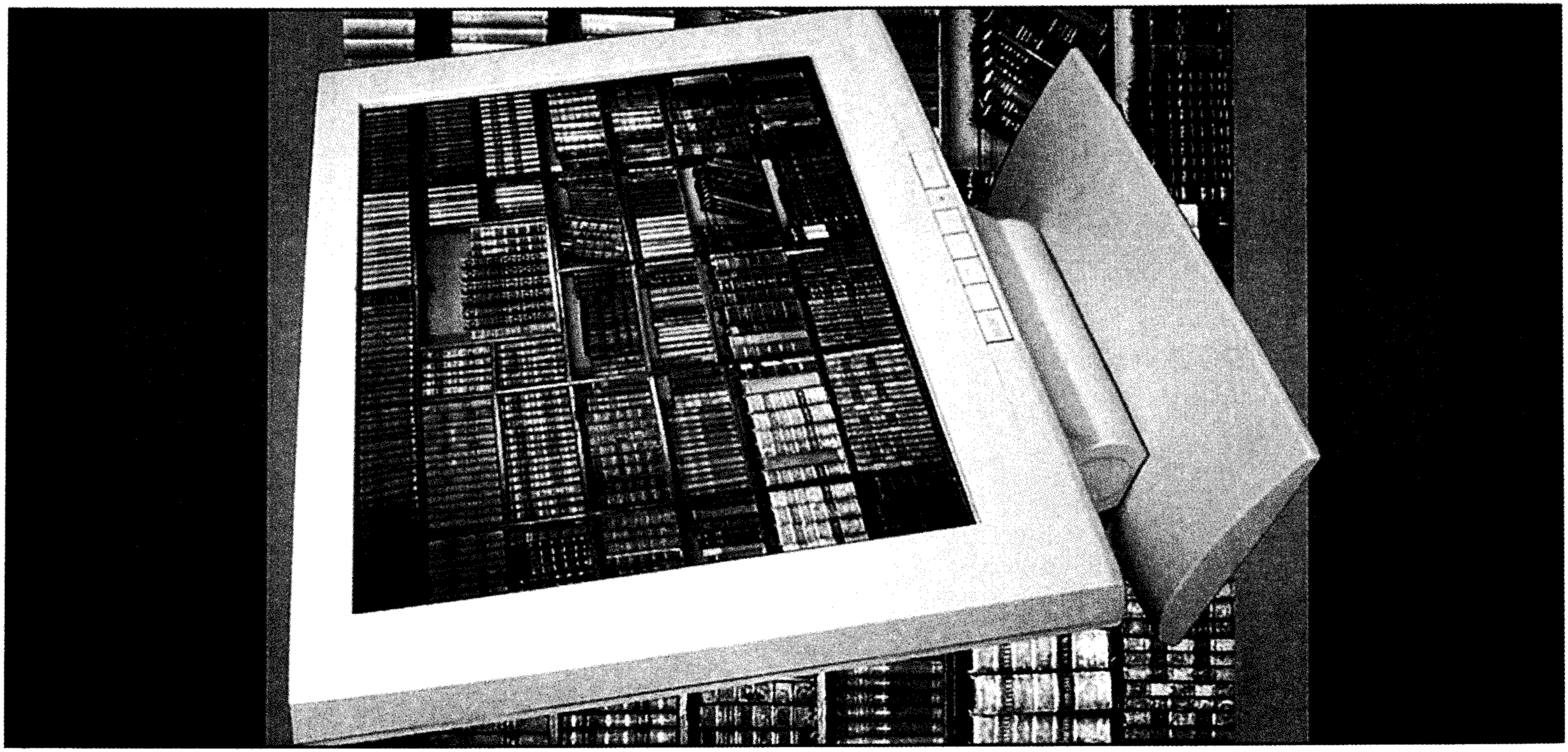
Silicon tetrafluoride will assume the geometric form of a tetrahedron, in much the same way as carbon and hydrogen form methane.

Their dipoles cancel out.

42. D

The strength of London dispersion forces, the only intermolecular attractions available to non-polar Cl_2 and I_2 molecules, is proportional to the total numbers of electrons within their molecules. With over three times as many electrons per molecule as chlorine, iodine molecules experience dispersion forces of greater strength than chlorine. This, in part, explains iodine's much higher boiling point than the boiling point of chlorine.

Appendices



CHEMISTRY 11 – DATA SHEETS

COMMON FORMULAS AND UNITS OF MEASUREMENT

- *STP conditions*: 0°C (273.15 K) and 1 atm (101.325 kPa) pressure.
- *SATP conditions*: 25°C (298.15 K) and 100 kPa pressure.
- *Molar Volumes*: $V_{\mathrm{STP}} = 22.4$ L and $V_{\mathrm{SATP}} = 24.8$ L
- 1 atm = 101.325 kPa = 760 mm Hg = 760 torr
- $T(\mathrm{K}) = T\,(^\circ\mathrm{C}) + 273.15$
- *Ideal Gas Law*: $PV = nRT$
- *Combined Gas Law*: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
- *The Gas Constant*: $R = 8.314\frac{\mathrm{L\cdot kPa}}{\mathrm{mol\cdot K}}$ or $0.082\,06\frac{\mathrm{L\cdot atm}}{\mathrm{mol\cdot K}}$
- *Avogadro's Number*: $N_{\mathrm{A}} = 6.022\times10^{23}$
- *Faraday constant:* $F = 9.65\times10^{4}$ C/mol e$^-$
- $K_{\mathrm{w}} = 1.0\times10^{-14}\left(\frac{\mathrm{mol}}{\mathrm{L}}\right)^2$ at 25° C
- $\mathrm{pH} = -\log\left[\mathrm{H}^+_{(aq)}\right]$ *note*: $\left(\left[\mathrm{H}^+_{(aq)}\right]\text{ in }\frac{\mathrm{mol}}{\mathrm{L}}\right)$
- $\left[\mathrm{H}^+_{(aq)}\right] = 10^{-\mathrm{pH}}$
- $C_1V_1 = C_2V_2$ (for dilution only)
- Water autoionization constant (Dissociation constant)
  $K_{\mathrm{w}} = 1.0\times10^{-14}$ at 298.15 K (for ion concentrations in mol/L)
- *Quadratic formula:* $x = \frac{-b\pm\sqrt{b^2-4ac}}{2a}$

SOLUBILITY OF SOME COMMON IONIC COMPOUNDS IN WATER AT 298.15 K (25°C)

ION	Group IA NH_4^+ $H^+(H_3O^+)$	ClO_3^- NO_3^- ClO_4^-	CH_3COO^-	Cl^- Br^- I^-	SO_4^{2-}	S^{2-}	OH^-	PO_4^{3-} SO_3^{2-} CO_3^{2-}
Solubility greater than or equal to 0.1 mol/L (very soluble)	all	all	most	most	most	Group IA Group IIA NH_4^+	Group IA NH_4^+ Sr^{2+} Ba^{2+} Tl^+	Group IA NH_4^+
Solubility less than 0.1 mol/L (slightly soluble)	none	none	Ag^+ Hg^+	Ag^+ Pb^{2+} Hg^+ Cu^+ Tl^+	Ca^{2+} Sr^{2+} Ba^{2+} Ra^{2+} Pb^{2+} Ag^+	most	most	most

COLOR CHART OF COMMON AQUEOUS IONS

ION	SYMBOL	COLOR
chromate	CrO_4^{2-}	yellow
chromium(III)	Cr^{3+}	green
chromium(II)	Cr^{2+}	blue
cobalt(II)	Co^{2+}	pink
copper(I)	Cu^+	green
copper(II)	Cu^{2+}	blue
dichromate	$Cr_2O_7^{2-}$	orange
iron(II)	Fe^{2+}	pale green
iron(III)	Fe^{3+}	pale yellow
manganese(II)	Mn^{2+}	pale pink
permanganate	MnO_4^-	purple

Specific Heat Capacity (at 298.15 K and 100.000 kPa)

$c_{\text{air}} = 1.01\ \text{J}/(\text{g}\cdot°\text{C})$

$c_{\text{polystyrene foam cup}} = 1.01\ \text{J}/(\text{g}\cdot°\text{C})$

$c_{\text{copper}} = 0.385\ \text{J}/(\text{g}\cdot°\text{C})$

$c_{\text{aluminium}} = 0.897\ \text{J}/(\text{g}\cdot°\text{C})$

$c_{\text{tin}} = 0.227\ \text{J}/(\text{g}\cdot°\text{C})$

$c_{\text{water}} = 4.19\ \text{J}/(\text{g}\cdot°\text{C})$

Periodic Table of the Elements (2006)

Symbol	**Au** 79	Atomic number
Element name	Gold	
	Atomic mass 196.96657	Atomic mass (g/mol)
	Atomic radius 135	Atomic radius (pm)

Alkali metals	Alkaline earth metals	Lanthanides	Actinides	Transition metals
Poor metals	Metalloids	Nonmetals	Halogens	Noble gases

Phase at standard temperature and pressure:
solid, liquid, gas

	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9	Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
Period 1	**H** 1 Hydrogen																	**He** 2 Helium
Period 2	**Li** 3 Lithium	**Be** 4 Beryllium											**B** 5 Boron	**C** 6 Carbon	**N** 7 Nitrogen	**O** 8 Oxygen	**F** 9 Fluorine	**Ne** 10 Neon
Period 3	**Na** 11 Sodium	**Mg** 12 Magnesium											**Al** 13 Aluminium	**Si** 14 Silicon	**P** 15 Phosphorus	**S** 16 Sulfur	**Cl** 17 Chlorine	**Ar** 18 Argon
Period 4	**K** 19 Potassium	**Ca** 20 Calcium	**Sc** 21 Scandium	**Ti** 22 Titanium	**V** 23 Vanadium	**Cr** 24 Chromium	**Mn** 25 Manganese	**Fe** 26 Iron	**Co** 27 Cobalt	**Ni** 28 Nickel	**Cu** 29 Copper	**Zn** 30 Zinc	**Ga** 31 Gallium	**Ge** 32 Germanium	**As** 33 Arsenic	**Se** 34 Selenium	**Br** 35 Bromine	**Kr** 36 Krypton
Period 5	**Rb** 37 Rubidium	**Sr** 38 Strontium	**Y** 39 Yttrium	**Zr** 40 Zirconium	**Nb** 41 Niobium	**Mo** 42 Molybdenum	**Tc** 43 Technetium	**Ru** 44 Ruthenium	**Rh** 45 Rhodium	**Pd** 46 Palladium	**Ag** 47 Silver	**Cd** 48 Cadmium	**In** 49 Indium	**Sn** 50 Tin	**Sb** 51 Antimony	**Te** 52 Tellurium	**I** 53 Iodine	**Xe** 54 Xenon
Period 6	**Cs** 55 Caesium	**Ba** 56 Barium		**Hf** 72 Hafnium	**Ta** 73 Tantalum	**W** 74 Tungsten	**Re** 75 Rhenium	**Os** 76 Osmium	**Ir** 77 Iridium	**Pt** 78 Platinum	**Au** 79 Gold	**Hg** 80 Mercury	**Tl** 81 Thallium	**Pb** 82 Lead	**Bi** 83 Bismuth	**Po** 84 Polonium	**At** 85 Astatine	**Rn** 86 Radon
Period 7	**Fr** 87 Francium	**Ra** 88 Radium		**Rf** 104 Rutherfordium	**Db** 105 Dubnium	**Sg** 106 Seaborgium	**Bh** 107 Bohrium	**Hs** 108 Hassium	**Mt** 109 Meitnerium	**Ds** 110 Darmstadtium	**Rg** 111 Roentgenium	**Uub** 112 Ununbium	**Uut** 113 Ununtrium	**Uuq** 114 Ununquadium	**Uup** 115 Ununpentium	**Uuh** 116 Ununhexium	**Uus** 117 Ununseptium	**Uuo** 118 Ununoctium

La 57 Lanthanum	**Ce** 58 Cerium	**Pr** 59 Praseodymium	**Nd** 60 Neodymium	**Pm** 61 Promethium	**Sm** 62 Samarium	**Eu** 63 Europium	**Gd** 64 Gadolinium	**Tb** 65 Terbium	**Dy** 66 Dysprosium	**Ho** 67 Holmium	**Er** 68 Erbium	**Tm** 69 Thulium	**Yb** 70 Ytterbium	**Lu** 71 Lutetium
Ac 89 Actinium	**Th** 90 Thorium	**Pa** 91 Protactinium	**U** 92 Uranium	**Np** 93 Neptunium	**Pu** 94 Plutonium	**Am** 95 Americium	**Cm** 96 Curium	**Bk** 97 Berkelium	**Cf** 98 Californium	**Es** 99 Einsteinium	**Fm** 100 Fermium	**Md** 101 Mendelevium	**No** 102 Nobelium	**Lr** 103 Lawrencium

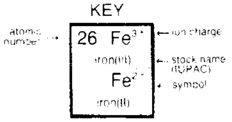
PERIODIC CHART OF IONS

TABLE OF POLYATOMIC IONS

acetate	CH_3COO^-	dichromate	$Cr_2O_7^{2-}$	dihydrogen phosphate	$H_2PO_4^-$
ammonium	NH_4^+	cyanide	CN^-	silicate	SiO_3^{2-}
benzoate	$C_6H_5COO^-$	hydroxide	OH^-	sulphate	SO_4^{2-}
borate	BO_3^{3-}	iodate	IO_3^-	sulphite	SO_3^{2-}
carbonate	CO_3^{2-}	nitrate	NO_3^-	hydrogen sulphide	HS^-
hydrogen carbonate	HCO_3^-	nitrite	NO_2^-	hydrogen sulphate	HSO_4^-
chlorate	ClO_3^-	oxalate	$OOCCOO^{2-}$	hydrogen sulphite	HSO_3^-
hypochlorite	ClO^-	permanganate	MnO_4^-	thiocyanate	SCN^-
chromate	CrO_4^{2-}	phosphate	PO_4^{3-}	thiosulphate	$S_2O_3^{2-}$
		hydrogen phosphate	HPO_4^{2-}		

1 H^+ hydrogen																1 H hydride	2 **He** **helium**
3 Li^+ lithium	4 Be^{2+} beryllium											5 **B** **boron**	6 **C** **carbon**	7 N^{3-} nitride	8 O^{2-} oxide	9 F fluoride	10 **Ne** **neon**
11 Na^+ sodium	12 Mg^{2+} magnesium											13 Al^{3+} aluminum	14 **Si** **silicon**	15 P^{3-} phosphide	16 S^{2-} sulphide	17 Cl chloride	18 **Ar** **argon**
19 K^+ potassium	20 Ca^{2+} calcium	21 Sc^{3+} scandium	22 Ti^{4+} titanium(IV) Ti^{3+} titanium(III)	23 V^{5+} vanadium(V) V^{4+} vanadium(IV)	24 Cr^{3+} chromium(III) Cr^{2+} chromium(II)	25 Mn^{2+} manganese(II) Mn^{4+} manganese(IV)	26 Fe^{3+} iron(III) Fe^{2+} iron(II)	27 Co^{2+} cobalt(II) Co^{3+} cobalt(III)	28 Ni^{2+} nickel(II) Ni^{3+} nickel(III)	29 Cu^{2+} copper(II) Cu^+ copper(I)	30 Zn^{2+} zinc	31 Ga^{3+} gallium	32 Ge^{4+} germanium	33 As^{3-} arsenide	34 Se^{2-} selenide	35 Br bromide	36 **Kr** **krypton**
37 Rb^+ rubidium	38 Sr^{2+} strontium	39 Y^{3+} yttrium	40 Zr^{4+} zirconium	41 Nb^{5+} niobium(V) Nb^{3+} niobium(III)	42 Mo^{6+} molybdenum	43 Tc^{7+} technetium	44 Ru^{3+} ruthenium(III) Ru^{4+} ruthenium(IV)	45 Rh^{3+} rhodium	46 Pd^{2+} palladium(II) Pd^{4+} palladium(IV)	47 Ag^+ silver	48 Cd^{2+} cadmium	49 In^{3+} indium	50 Sn^{4+} tin(IV) Sn^{2+} tin(II)	51 Sb^{3+} antimony(III) Sb^{5+} antimony(V)	52 Te^{2-} telluride	53 I iodide	54 **Xe** **xenon**
55 Cs^+ cesium	56 Ba^{2+} barium	57-71	72 Hf^{4+} hafnium	73 Ta^{5+} tantalum	74 W^{6+} tungsten	75 Re^{7+} rhenium	76 Os^{4+} osmium	77 Ir^{4+} iridium	78 Pt^{4+} platinum(IV) Pt^{2+} platinum(II)	79 Au^{3+} gold(III) Au^+ gold(I)	80 Hg^{2+} mercury(II) Hg^+ mercury(I)	81 Tl^+ thallium(I) Tl^{3+} thallium(III)	82 Pb^{2+} lead(II) Pb^{4+} lead(IV)	83 Bi^{3+} bismuth(III) Bi^{5+} bismuth(V)	84 Po^{2+} polonium(II) Po^{4+} polonium(IV)	85 At astatide	86 **Rn** **radon**
87 Fr^+ francium	88 Ra^{2+} radium	89-103															

KEY



57 La^{3+} lanthanum	58 Ce^{3+} cerium	59 Pr^{3+} praseodymium	60 Nd^{3+} neodymium	61 Pm^{3+} promethium	62 Sm^{3+} samarium(III) Sm^{2+} samarium(II)	63 Eu^{3+} europium(III) Eu^{2+} europium(II)	64 Gd^{3+} gadolinium	65 Tb^{3+} terbium	66 Dy^{3+} dysprosium	67 Ho^{3+} holmium	68 Er^{3+} erbium	69 Tm^{3+} thulium	70 Yb^{3+} ytterbium(III) Yb^{2+} ytterbium(II)	71 Lu^{2+} lutetium
89 Ac^{3+} actinium	90 Th^{4+} thorium	91 Pa^{5+} protactinium(V) Pa^{4+} protactinium(IV)	92 U^{6+} uranium(VI) U^{4+} uranium(IV)	93 Np^{5+} neptunium	94 Pu^{4+} plutonium(IV) Pu^{6+} plutonium(VI)	95 Am^{3+} americium(III) Am^{4+} americium(IV)	96 Cm^{3+} curium	97 Bk^{3+} berkelium(III) Bk^{4+} berkelium(IV)	98 Cf^{3+} californium	99 Es^{3+} einsteinium	100 Fm^{3+} fermium	101 Md^{2+} mendelevium(II) Md^{3+} mendelevium(III)	102 No^{2+} nobelium(II) No^{3+} nobelium(III)	103 Lr^{3+} lawrencium

PERIODIC TABLE OF ELECTRONEGATIVITY USING THE PAULING SCALE

→ Atomic radius decreases → Ionization energy increases → Electronegativity increases →

Group (vertical) **Period** (horizontal)	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 2.20																	He
2	Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	[illegible]	[illegible]	Ne
3	Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
5	Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.67
6	Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 2.33	Bi 2.02	Po 2.0	At 2.2	Rn
7	Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo

Lanthanides	*	La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27
Actinides	**	Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr

Selected SI Prefixes

Prefix	Exponential Symbol	Value
tera	T	10^{12}
giga	G	10^{9}
mega	M	10^{6}
kilo	k	10^{3}
milli	m	10^{-3}
micro	µ	10^{-6}
nano	n	10^{-9}
pico	p	10^{-12}

Table of Selected Standard Electrode Potentials*

Reduction Half-Reaction	Electrical Potential $E°$ (V)
$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$	+2.87
$PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons PbSO_4(s) + 2H_2O(l)$	+1.69
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$Au^{3+}(aq) + 3e^- \rightleftharpoons Au(s)$	+1.50
$ClO_4^-(aq) + 8H^+(aq) + 8e^- \rightleftharpoons Cl^-(aq) + 4H_2O(l)$	+1.39
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$2HNO_2(aq) + 4H^+(aq) + 4e^- \rightleftharpoons N_2O(g) + 3H_2O(l)$	+1.30
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.23
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
$MnO_2(s) + 4H^+(aq) + 2e^- \rightleftharpoons Mn^{2+}(aq) + 2H_2O(l)$	+1.22
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.07
$Hg^{2+}(aq) + 2e^- \rightleftharpoons Hg(l)$	+0.85
$OCl^-(aq) + H_2O(l) + 2e^- \rightleftharpoons Cl^-(aq) + 2OH^-(aq)$	+0.84
$2NO_3^-(aq) + 4H^+(aq) + 2e^- \rightleftharpoons N_2O_4(g) + 2H_2O(l)$	+0.80
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(l)$	+0.70
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons H_2SO_3(aq) + H_2O(l)$	+0.17
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(aq)$	+0.14
$AgBr(s) + e^- \rightleftharpoons Ag(s) + Br^-(aq)$	+0.07
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14
$AgI(s) + e^- \rightleftharpoons Ag(s) + I^-(aq)$	−0.15
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	−0.26
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	−0.28
$PbSO_4(s) + 2e^- \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	−0.36
$Se(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2Se(aq)$	−0.40
$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$	−0.40
$Cr^{3+}(aq) + e^- \rightleftharpoons Cr^{2+}(aq)$	−0.41
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.45
$NO_2^-(aq) + H_2O(l) + e^- \rightleftharpoons NO(g) + 2OH^-(aq)$	−0.46
$Ag_2S(s) + 2e^- \rightleftharpoons 2Ag(s) + S^{2-}(aq)$	−0.69
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	−0.83
$Cr^{2+}(aq) + 2e^- \rightleftharpoons Cr(s)$	−0.91
$Se(s) + 2e^- \rightleftharpoons Se^{2-}(aq)$	−0.92
$SO_4^{2-}(aq) + H_2O(l) + 2e^- \rightleftharpoons SO_3^{2-}(aq) + 2OH^-(aq)$	−0.93
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	−1.66
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	−2.37
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$	−2.87
$Ba^{2+}(aq) + 2e^- \rightleftharpoons Ba(s)$	−2.91
$K^+(aq) + e^- \rightleftharpoons K(s)$	−2.93
$Li^+(aq) + e^- \rightleftharpoons Li(s)$	−3.04

*For 1.0 mol/L solutions at 298.15 K (25.00 °C) and a pressure of 101.325 kPa

ACID BASE INDICATOR CHART

Indicator	pH Range	Quantity per 10 ml	Acid	Base
Thymol Blue	1.2-2.8	1–2 drops 0.1% soln. in aq.	red	yellow
Pentamethoxy red	1.2-2.3	1 drop 0.1% soln. in 70% alc.	red-violet	colorless
Tropeolin OO	1.3-3.2	1 drop 1% aq. soln.	red	yellow
2,4-Dinitrophenol	2.4-4.0	1–2 drops 0.1% soln. in 50% alc.	colorless	yellow
Methyl yellow	2.9–4.0	1 drop 0.1% soln. in 90% alc.	red	yellow
Methyl orange	3.1–4.4	1 drop 0.1% aq. soln.	red	orange
Bromphenol blue	3.0–4.6	1 drop 0.1% aq. soln.	yellow	blue violet
Tetrabromphenol blue	3.0–4.6	1 drop 0.1% aq. soln.	yellow	blue
Alizarin sodium sulfonate	3.7–5.2	1 drop 0.1% aq. soln.	yellow	violet
α–Naphthyl red	3.7–5.0	1 drop 0.1% soln. in 70% alc.	red	yellow
p–Ethoxychrysoidine	3.5–5.5	1 drop 0.1% aq. soln.	red	yellow
Bromcresol green	4.0–5.6	1 drop 0.1% aq. soln.	yellow	blue
Methyl red	4.4–6.2	1 drop 0.1% aq. soln.	red	yellow
Bromcresol purple	5.2–6.8	1 drop 0.1% aq. soln.	yellow	purple
Chlorphenol red	5.4–6.8	1 drop 0.1% aq. soln.	yellow	red
Bromphenol blue	6.2–7.6	1 drop 0.1% aq. soln.	yellow	blue
p–Nitrophenol	5.0–7.0	1–5 drops 0.1% aq. soln.	colorless	yellow
Azolitmin	5.0–8.0	5 drops 0.5% aq. soln.	red	blue
Phenol red	6.4–8.0	1 drop 0.1% aq. soln.	yellow	red
Neutral red	6.8–8.0	1 drop 0.1% soln. in 70% alc.	red	yellow
Rosolic acid	6.8–8.0	1 drop 0.1% soln. in 90% alc.	yellow	red
Cresol red	7.2–8.8	1 drop 0.1% aq. soln.	yellow	red
α–Naphtholphthalein	7.3–8.7	1–5 drops 0.1% soln. in 70% alc.	rose	green
Tropeolin OOO	7.6–8.9	1 drop 0.1% aq. soln.	yellow	rose–red
Thymol blue	8.0–9.6	1–5 drops 0.1% aq. soln.	yellow	blue
Phenolphthalein	8.0–10.0	1–5 drops 0.1% soln. in 70% alc.	colorless	red
α–Naphtholbenzein	9.0–11.0	1–5 drops 0.1% soln. in 90% alc.	yellow	blue
Thymolphthalein	9.4–10.6	1 drop 0.1% soln. in 90% alc.	colorless	blue
Nile blue	10.1–11.1	1 drop 0.1% aq. soln.	blue	red
Alizarin yellow	10.0–12.0	1 drop 0.1% aq. soln.	yellow	lilac
Salicyl yellow	10.0–12.0	1–5 drops 0.1% soln. in 90% alc.	yellow	orange brown
Diazo violet	10.1–12.0	1 drop 0.1% aq. soln.	yellow	violet
Tropeolin O	11.0–13.0	1 drop 0.1% aq. soln.	yellow	orange–brown
Nitramine	11.0–13.0	1–2 drops 0.1% soln in 70% alc.	colorless	orange brown
Poirrier's blue	11.0–13.0	1 drop 0.1% aq. soln.	blue	violet–pink
Trinitrobenzoic acid	12.0–13.4	1 drop 0.1% aq. soln.	colorless	orange red

Standard Molar Enthalpies of Formation at 298.15 K

Name	Formula	$\Delta_f H^\circ$ (kJ/mol)
aluminum oxide	$Al_2O_3(s)$	−1 675.7
ammonia	$NH_3(g)$	−45.9
ammonium chloride	$NH_4Cl(s)$	−314.4
ammonium nitrate	$NH_4NO_3(s)$	−365.6
barium carbonate	$BaCO_3(s)$	−1 213.0
barium chloride	$BaCl_2(s)$	−855.0
barium hydroxide	$Ba(OH)_2(s)$	−944.7
barium oxide	$BaO(s)$	−548.0
barium sulfate	$BaSO_4(s)$	−1 473.2
benzene	$C_6H_6(l)$	+49.1
butane	$C_4H_{10}(g)$	−125.7
calcium carbonate	$CaCO_3(s)$	−1 207.6
calcium chloride	$CaCl_2(s)$	−795.4
calcium hydroxide	$Ca(OH)_2(s)$	−985.2
calcium oxide	$CaO(s)$	−634.9
calcium sulfate	$CaSO_4(s)$	−1 434.5
carbon dioxide	$CO_2(g)$	−393.5
carbon monoxide	$CO(g)$	−110.5
chromium(III) oxide	$Cr_2O_3(s)$	−1 139.7
copper(I) oxide	$Cu_2O(s)$	−168.6
copper(II) oxide	$CuO(s)$	−157.3
copper(II) sulfate	$CuSO_4(s)$	−771.4
copper(I) sulfide	$Cu_2S(s)$	−79.5
copper(II) sulfide	$CuS(s)$	−53.1
dinitrogen tetroxide	$N_2O_4(g)$	+11.1
ethane	$C_2H_6(g)$	−84.0
ethanoic acid (acetic acid)	$CH_3COOH(l)$	−484.3
ethanol	$C_2H_5OH(l)$	−277.6
ethene (ethylene)	$C_2H_4(g)$	+52.4
ethyne (acetylene)	$C_2H_2(g)$	+227.4
glucose	$C_6H_{12}O_6(s)$	−1 273.3
hydrogen bromide	$HBr(g)$	−36.3
hydrogen chloride	$HCl(g)$	−92.3
hydrogen fluoride	$HF(g)$	−273.3
hydrogen iodide	$HI(g)$	+26.5
hydrogen perchlorate	$HClO_4(l)$	−40.6
hydrogen peroxide	$H_2O_2(l)$	−187.8
hydrogen sulfide	$H_2S(g)$	−20.6
iron(II) oxide	$FeO(s)$	−272.0
iron(III) oxide	$Fe_2O_3(s)$	−824.2
iron(II,III) oxide (magnetite)	$Fe_3O_4(s)$	−1 118.4
lead(II) bromide	$PbBr_2(s)$	−278.7
lead(II) chloride	$PbCl_2(s)$	−359.4
lead(II) oxide (red)	$PbO(s)$	−219.0
lead(IV) oxide	$PbO_2(s)$	−277.4
magnesium carbonate	$MgCO_3(s)$	−1 095.8
magnesium chloride	$MgCl_2(s)$	−641.3

Standard Molar Enthalpies of Formation at 298.15 K cont'd

Name	Formula	$\Delta_f H^\circ$ (kJ/mol)
magnesium hydroxide	$Mg(OH)_2(s)$	-924.5
magnesium oxide	$MgO(s)$	-601.6
magnesium sulfate	$MgSO_4(s)$	$-1\ 284.9$
manganese(II) oxide	$MnO(s)$	-385.2
manganese(IV) oxide	$MnO_2(s)$	-520.0
mercury(II) oxide (red)	$HgO(s)$	-90.8
mercury(II) sulfide (red)	$HgS(s)$	-58.2
methanal (formaldehyde)	$CH_2O(g)$	-108.6
methane	$CH_4(g)$	-74.6
methanoic acid (formic acid)	$HCOOH(l)$	-425.0
methanol	$CH_3OH(l)$	-239.2
nickel(II) oxide	$NiO(s)$	-240.6
nitric acid	$HNO_3(l)$	-174.1
nitrogen dioxide	$NO_2(g)$	$+33.2$
nitrogen monoxide	$NO(g)$	$+91.3$
octane	$C_8H_{18}(l)$	-250.1
pentane	$C_5H_{12}(l)$	-173.5
phosphorus pentachloride	$PCl_5(s)$	-443.5
phosphorus trichloride (liquid)	$PCl_3(l)$	-319.7
phosphorus trichloride (vapour)	$PCl_3(g)$	-287.0
potassium bromide	$KBr(s)$	-393.8
potassium chlorate	$KClO_3(s)$	-397.7
potassium chloride	$KCl(s)$	-436.5
potassium hydroxide	$KOH(s)$	-424.6
propane	$C_3H_8(g)$	-103.8
silicon dioxide (α-quartz)	$SiO_2(s)$	-910.7
silver bromide	$AgBr(s)$	-100.4
silver chloride	$AgCl(s)$	-127.0
silver iodide	$AgI(s)$	-61.8
sodium bromide	$NaBr(s)$	-361.1
sodium chloride	$NaCl(s)$	-411.2
sodium hydroxide	$NaOH(s)$	-425.8
sodium iodide	$NaI(s)$	-287.8
sucrose	$C_{12}H_{22}O_{11}(s)$	$-2\ 226.1$
sulfur dioxide	$SO_2(g)$	-296.8
sulfuric acid	$H_2SO_4(l)$	-814.0
sulfur trioxide (liquid)	$SO_3(l)$	-441.0
sulfur trioxide (vapour)	$SO_3(g)$	-395.7
tin(II) chloride	$SnCl_2(s)$	-325.1
tin(IV) chloride	$SnCl_4(l)$	-511.3
tin(II) oxide	$SnO(s)$	-280.7
tin(IV) oxide	$SnO_2(s)$	-577.6
water (liquid)	$H_2O(l)$	-285.8
water (vapour)	$H_2O(g)$	-241.8
zinc oxide	$ZnO(s)$	-350.5
zinc sulfide (sphalerite)	$ZnS(s)$	-206.0

Relative Strengths of Acids and Bases at 298.15 K

Common Name **IUPAC / Systematic Name**	**Acid Formula**	**Conjugate Base Formula**	K_a
perchloric acid aqueous hydrogen perchlorate	$HClO_4(aq)$	$ClO_4^-(aq)$	very large
hydroiodic acid aqueous hydrogen iodide	$HI(aq)$	$I^-(aq)$	very large
hydrobromic acid aqueous hydrogen bromide	$HBr(aq)$	$Br^-(aq)$	very large
hydrochloric acid aqueous hydrogen chloride	$HCl(aq)$	$Cl^-(aq)$	very large
sulfuric acid aqueous hydrogen sulfate	$H_2SO_4(aq)$	$HSO_4^-(aq)$	very large
nitric acid aqueous hydrogen nitrate	$HNO_3(aq)$	$NO_3^-(aq)$	very large
hydronium ion	$H_3O^+(aq)$	$H_2O(l)$	1
oxalic acid	$HOOCCOOH(aq)$	$HOOCCOO^-(aq)$	5.6×10^{-2}
sulfurous acid aqueous hydrogen sulfite	$H_2SO_3(aq)$	$HSO_3^-(aq)$	1.4×10^{-2}
hydrogen sulfate ion	$HSO_4^-(aq)$	$SO_4^{2-}(aq)$	1.0×10^{-2}
phosphoric acid aqueous hydrogen phosphate	$H_3PO_4(aq)$	$H_2PO_4^-(aq)$	6.9×10^{-3}
citric acid 2-hydroxy-1,2,3-propanetricarboxylic acid	$C_3H_5O(COOH)_3(aq)$	$C_3H_5O(COOH)_2COO^-(aq)$	7.4×10^{-4}
hydrofluoric acid aqueous hydrogen fluoride	$HF(aq)$	$F^-(aq)$	6.3×10^{-4}
nitrous acid aqueous hydrogen nitrite	$HNO_2(aq)$	$NO_2^-(aq)$	5.6×10^{-4}
formic acid methanoic acid	$HCOOH(aq)$	$HCOO^-(aq)$	1.8×10^{-4}
hydrogen oxalate ion	$HOOCCOO^-(aq)$	$OOCCOO^{2-}(aq)$	1.5×10^{-4}
lactic acid 2-hydroxypropanoic acid	$C_2H_5OCOOH(aq)$	$C_2H_5OCOO^-(aq)$	1.4×10^{-4}
ascorbic acid 2(1,2-dihydroxyethyl)-4,5-dihydroxy -furan-3-one	$H_2C_6H_6O_6(aq)$	$HC_6H_6O_6^-(aq)$	9.1×10^{-5}
benzoic acid benzenecarboxylic acid	$C_6H_5COOH(aq)$	$C_6H_5COO^-(aq)$	6.3×10^{-5}
acetic acid ethanoic acid	$CH_3COOH(aq)$	$CH_3COO^-(aq)$	1.8×10^{-5}

Relative Strengths of Acids and Bases at 298.15 K

Common Name **IUPAC / Systematic Name**	**Acid Formula**	**Conjugate Base Formula**	K_a
dihydrogen citrate ion	$C_3H_5O(COOH)_2COO^-(aq)$	$C_3H_5OCOOH(COO)_2^{2-}(aq)$	1.7×10^{-5}
butanoic acid	$C_3H_7COOH(aq)$	$C_3H_7COO^-(aq)$	1.5×10^{-5}
propanoic acid	$C_2H_5COOH(aq)$	$C_2H_5COO^-(aq)$	1.3×10^{-5}
carbonic acid ($CO_2 + H_2O$) aqueous hydrogen carbonate	$H_2CO_3(aq)$	$HCO_3^-(aq)$	4.5×10^{-7}
hydrogen citrate ion	$C_3H_5OCOOH(COO)_2^{2-}(aq)$	$C_3H_5O(COO)_3^{3-}(aq)$	4.0×10^{-7}
hydrosulfuric acid aqueous hydrogen sulfide	$H_2S(aq)$	$HS^-(aq)$	8.9×10^{-8}
hydrogen sulfite ion	$HSO_3^-(aq)$	$SO_3^{2-}(aq)$	6.3×10^{-8}
dihydrogen phosphate ion	$H_2PO_4^-(aq)$	$HPO_4^{2-}(aq)$	6.2×10^{-8}
hypochlorous acid aqueous hydrogen hypochlorite	$HOCl(aq)$	$OCl^-(aq)$	4.0×10^{-8}
hydrocyanic acid aqueous hydrogen cyanide	$HCN(aq)$	$CN^-(aq)$	6.2×10^{-10}
ammonium ion	$NH_4^+(aq)$	$NH_3(aq)$	5.6×10^{-10}
hydrogen carbonate ion	$HCO_3^-(aq)$	$CO_3^{2-}(aq)$	4.7×10^{-11}
hydrogen ascorbate ion	$HC_6H_6O_6^-(aq)$	$C_6H_6O_6^{2-}(aq)$	2.0×10^{-12}
hydrogen phosphate ion	$HPO_4^{2-}(aq)$	$PO_4^{3-}(aq)$	4.8×10^{-13}
water	$H_2O(l)$	$OH^-(aq)$	1.0×10^{-14}

Note: An approximation may be used instead of the quadratic formula when the concentration of H_3O^+ produced is less than 5% of the original acid concentration (or the concentration of the acid is 1 000 times greater than the K_a). The same approximation can also be used for weak bases. The formula of the carboxylic acids have been written so that the COOH group can be easily recognized. Either the common or IUPAC name is acceptable.

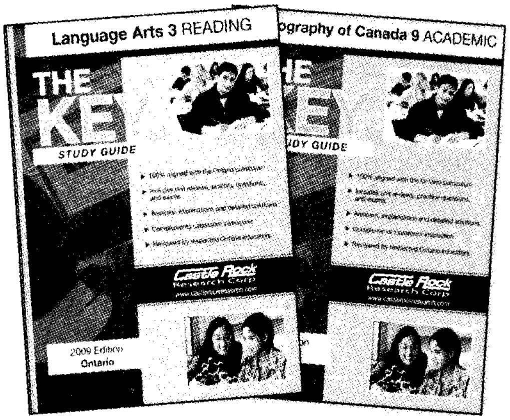
Colours of Common Aqueous Ions

Ionic Species	Solution Concentration	
	1.0 mol/L	0.010 mol/L
chromate	yellow	pale yellow
chromium(III)	blue-green	green
chromium(II)	dark blue	pale blue
cobalt(II)	red	pink
copper(I)	blue-green	pale blue-green
copper(II)	blue	pale blue
dichromate	orange	pale orange
iron(II)	lime green	colourless
iron(III)	orange-yellow	pale yellow
manganese(II)	pale pink	colourless
nickel(II)	blue-green	pale blue-green
permanganate	deep purple	purple-pink

Flame Colour of Elements

Element	Symbol	Colour
lithium	Li	red
sodium	Na	yellow
potassium	K	violet
rubidium	Rb	violet
cesium	Cs	violet
calcium	Ca	yellowish red
strontium	Sr	scarlet red
barium	Ba	yellowish green
copper	Cu	blue to green
boron	B	yellowish green
lead	Pb	blue-white

Note: The flame test can be used to determine the identity of a metal or a metal ion. Blue to green indicates a range of colours that might appear.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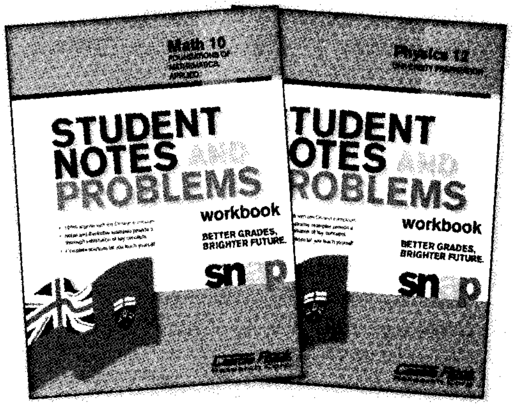
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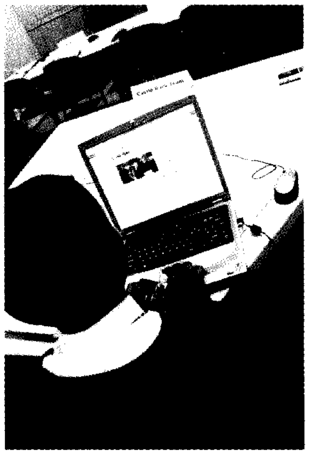

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