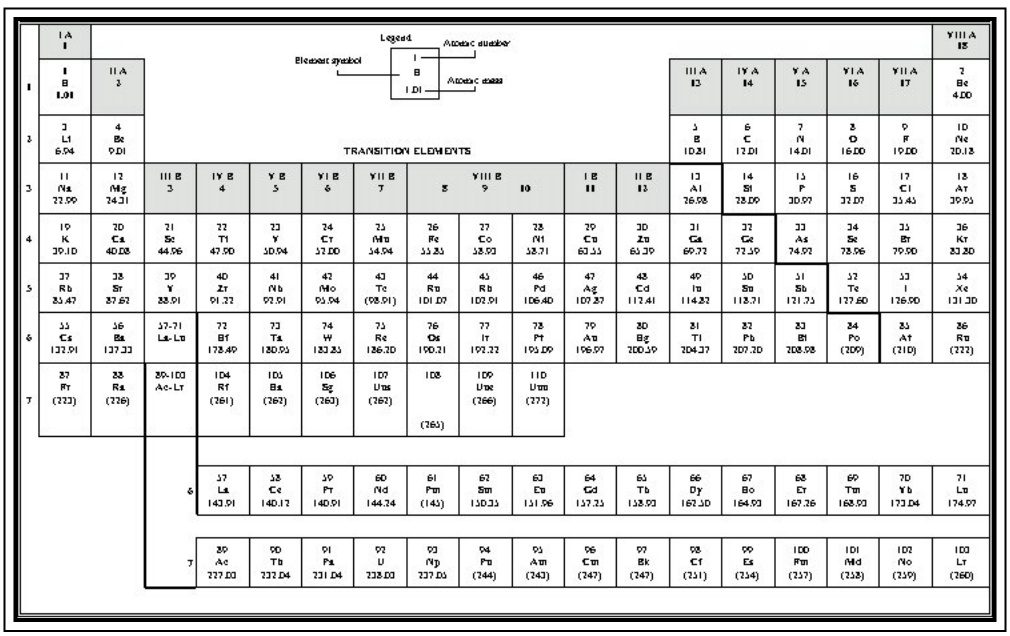
Student Study Guide Chemistry 534

(Revised Edition 2001)



Student Study Guide
Chemistry 534

Introduction

This Study Guide was written by a committee of Chemistry teachers to help students to prepare for the Chemistry 534 theory/written examination.

The contents of the Guide are:

Module 2	**Gases and their Applications**
Module 3	**Energy in Chemical Reactions**
Module 4	**Rate of Chemical Reactions**
Module 5	**Equilibrium in Chemical Reactions**
Appendices	**Formulas, Standard Reduction Potentials and Periodic Table**

Each module is presented in sections which cover one or more objectives of the program. These sections include *Key Concepts*, *Examples* and *Sample Questions*. The answers and solutions to the *Sample Questions* may be found at the end of the Guide. Objectives have occasionally been grouped together or presented in an order different from the MEQ program at the discretion of the authors.

The Guide is designed as a study tool for students and is not to be considered as an official course program. For a more detailed description of the course content and learning activities, please refer to the document 16-3177A, Secondary School Curriculum *Chemistry 534*: *The Discovery of Matter and Energy*, published by the Gouvernement du Québec, Ministère de l'Éducation 1992 ISBN: 2-550-23347-6.

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Feedback needed from Teachers!

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The content is adequate for

Module 2	YES ☐ NO ☐	**Module 3**	YES ☐ NO ☐
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I would suggest the following changes:

Module 2: ____

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Thank you for your assistance!

MODULE 2

- GASES AND THEIR APPLICATIONS -

1.1 & 1.2 The use of gases and how they relate to our quality of life, society and environment.

Gases play a very important role in our daily existence. Since we are surrounded by an ocean of gas called the atmosphere, many of the properties of gases are already familiar to us. We know that we can squeeze a balloon into a smaller shape and that perfume released into the corner of a room can, in time, be detected all over the room even if the air is still.

Gases such as carbon monoxide (CO), sulfur oxides (SO_2 and SO_3) and nitrogen oxides (NO and NO_2) have always been present in the atmosphere along with nitrogen gas (N_2), oxygen gas (O_2), and carbon dioxide (CO_2). Volcanic action produces sulfur oxides and lightning produces nitrogen oxides from atmospheric nitrogen and oxygen. In many areas of the world human activity has resulted in high concentrations of these oxides which has had adverse effects on human, animal and plant life.

Gases can be both harmful and beneficial. Chlorine, for example, has been used in the production of war materials and is more commonly used today in the manufacture of cleaning and disinfectant products. All modern cars are equipped with an inflatable air bag to protect passengers from harmful collisions. The gas which inflates the air bag is nitrogen, a common gas which is essential to life.

To combat photochemical smog, catalytic converters are attached to automobile exhaust systems to transform pollutants such as carbon monoxide (CO), and nitrogen monoxide (NO), into less harmful carbon dioxide (CO_2), and nitrogen gas (N_2).

Ozone (O_3), is a gas formed when some sort of electrical discharge passes through molecular oxygen (O_2). Ozone has a distinct odour that can be observed when using an electrical motor or photocopier or during an electrical storm. Ozone gas performs a vital function in the upper atmosphere by preventing harmful ultraviolet radiation from reaching the Earth.

Air conditioning increase our comfort in hot climatic conditions and one of the gases that has been used for cooling is freon (a chlorinated fluorocarbon or CFC). This gas is effective for cooling but because it attacks the ozone layer, it has been replaced by several less harmful products.

Some other gases with which you may be familiar and their uses are listed below:

- Ammonia gas (NH_3) is used as a refrigerant, in solution as a cleaning product and to make many nitrogen containing compounds.
- Helium gas is used to inflate balloons.
- Neon gas is used in lighting because of the bright red colour produced in electric discharge tubes.
- Propane (C_3H_8) is used to heat homes, cook food and fuel cars.
- Methane (CH_4) is also used as a fuel.

SAMPLE QUESTION

1. *Gases are used for different purposes such as:*

   1) Disinfection   2) Air conditioning   3) Energy production
   In order, which gases serve the purposes listed above?

A)	***1. Chlorine***	***2. Freon***	***3. Methane***
B)	***1. Nitrogen***	***2. Oxygen***	***3. Chlorine***
C)	***1. Methane***	***2. Nitrogen***	***3. Carbon dioxide***
D)	***1. Oxygen***	***2. Freon***	***3. Carbon dioxide***

2.1, 2.2 & 2.3 Factors that affect the volume of a gas

PRESSURE

KEY CONCEPTS

1. The S.I. unit for gas pressure is the kilopascal.
2. The Standard Pressure is defined to be 101.3 kilopascals (written 101.3 kPa).
3. Gas pressure is measured using an instrument called a manometer.
4. The basic relationship between Pressure and Volume is BOYLE'S LAW, which states that pressure and volume are inversely proportional. This is also called an inverse variation function in mathematics. This is expressed mathematically as either $P \cdot V = k$ where k is a constant, or more conveniently as: $P_1V_1 = P_2V_2$.

In general, the term pressure is defined as force per unit area. Gases exert a pressure on any surface with which they are in contact. A simple example of pressure would be a gas inside an inflated balloon where the gas exerts a pressure on the inside surface of the balloon.

An instrument called a **barometer** can measure atmospheric pressure and the measurement is reported in kilopascals (kPa).

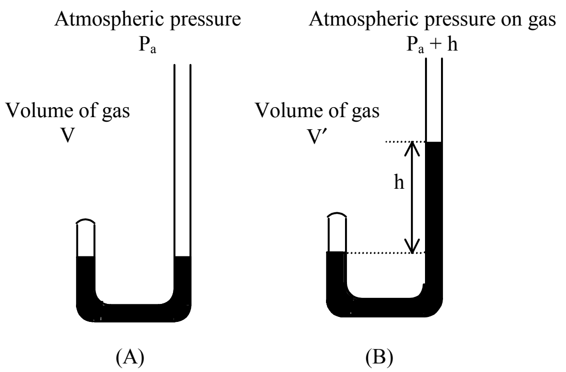
On pascal is defined to be the pressure exerted by a force of one newton acting on an area of one square metre. Using symbols, $1\ Pa = 1\ N/m^2$ and $1\ kPa = 1000\ Pa$.

Standard Atmospheric Pressure corresponds to the typical pressure found at sea level, which is the pressure sufficient to support a column of liquid mercury to a height of 760 mm. Converting this pressure to pascals it becomes 1.01325×10^5 Pa or 1.01325×10^2 kPa. *This is most commonly reported as* **101.3 kPa**.

In laboratories we often used a simple device, called a **manometer** to measure the pressure of gases in a container.

PRESSURE – VOLUME RELATIONSHIP

Robert Boyle was the first person to investigate the relationship between pressure and volume. He used a J-shaped tube like that shown in the accompanying figure: Diagram illustrating Boyle's Experiment.



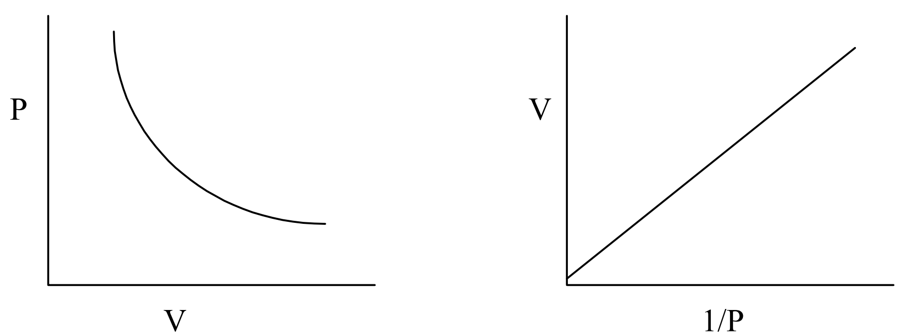
Boyle used a J-tube to study volume and pressure. One end of the J-tube is closed , trapping air at the end of the tube, (A). When the height of the mercury is the same in the open and the closed parts of the tube, the pressure exerted on the gas is equal to the atmospheric pressure, (B). The pressure of the gas is increased by adding mercury to the tube. The pressure exerted on the gas is equal to h (the difference in the heights of the two mercury surfaces) plus the atmospheric pressure, and the volume of the gas is smaller.

In his experiment, he trapped a quantity of gas above the mercury in the closed end of a J-shaped tube. Boyle varied the pressure in the tube by varying the amount of mercury. He found that the volume of the gas decreased as the pressure increased. For example, if he doubled the pressure he found that the volume of the gas was reduced to one-half of the original value.

This led to the relationship: *Pressure x Volume = constant*
Mathematically, this is written: $P \cdot V = k$

The value of the constant depends on the temperature and the amount of gas used in the sample. We can see from the accompanying diagram that a curve is always obtained for a given quantity of gas at a fixed temperature. If we plot the volume of the gas, V, against the reciprocal of P, 1/P, we obtain a linear relationship as can be seen in the following graphs.

Graphs based on Boyle's Law



Example:

A chemist collects 20.0 mL of a gas at a pressure of 250 kPa. What will be the volume of this sample of gas if the pressure is increased to 400 kPa?

Solution

Using the basic equation $P_1V_1 = P_2V_2$:
P_1 (starting pressure) $= 250$ kPa
V_1 (starting volume) $= 20.0$ mL
P_2 (final pressure) $= 400$ kPa
V_2 (final volume) = Unknown

$P_1 \cdot V_1 = P_2 \cdot V_2$
$V_2 = (P_1 \cdot V_1)/P_2$
$V_2 = (250 \text{ kPa})(20.0 \text{ mL})/400 \text{ kPa}$
$V_2 = 12.5$ mL

TEMPERATURE

KEY CONCEPTS

1. All problems involving gases must be solved using KELVIN measurement.

2. The conversion between Celsius and Kelvin measurement is:
   Celsius Temperature + 273 = Kelvin Temperature

3. The basic relationship between volume and Kelvin temperature is called Charles' Law which states that volume is directly proportional to Kelvin (or absolute) temperature. Mathematically, this is called a direct variation function and is given by $V = k \cdot T$ or more conveniently, as $V_1/T_1 = V_2/T_2$

4. Standard Temperature is 0^{o}C which must be converted to the Kelvin scale:
   $0^{o}\text{C} + 273 = 273$ K

Example:

A chemist discovers that 30.0 mL of a gas measured at 10.0^{o}C will increase in volume if the temperature is raised to 50.0^{o}C. What will be the corresponding volume for this increase in temperature?

Solution:

Use the relationship: $V_1/T_1 = V_2/T_2$

Remember to convert the temperature to the Kelvin scale.

V_1 (starting volume)	= 30.0 mL
T_1 (starting temperature)	= 10.0^oC + 273 = 283 K
T_2 (final temperature)	= 50.0^oC + 273 = 323 K
V_2 (final temperature)	= unknown

$V_1/T_1 = V_2/T_2$

$V_2 = (V_1T_2)/T_1$

$V_2 = (30.0\text{ mL})(323\text{ K})/(283\text{ K})$

$V_2 = 34.2\text{ mL}$

SAMPLE QUESTIONS

1. *What volume would 38.0 mL of air occupy when its pressure changes from 120 kPa to 95.0 kPa? (Assume constant temperature.)*

2. *80.0 mL of a gas was measured at a barometer reading of 102.4 kPa. The next day the barometer read 100.7 kPa. What was the new volume at this pressure? (Assume constant temperature.)*

3. *A gas has a volume of 8.00 L at a temperature of 25.0^{o}C. What will be the volume of the gas if the temperature is raised to 50.0^{o}C? (Assume the pressure is constant.)*

4. *Methane gas, CH_4, is collected in an experiment. If a student collects 12 mL of this gas at a temperature of 115^{o}C, what would be the Celsius temperature if this sample is cooled to a volume of 9.0 mL? (Assume constant pressure.)*

5. *If a student collects 52.0 mL of hydrogen gas at a temperature of 20.0^{o}C, what volume would the gas occupy at 28.0^{o}C if the pressure remains constant?*

2.5 The relationship between the behaviors of gases under different conditions.

UNIVERSAL GAS LAW PROBLEMS:

KEY CONCEPTS

1. When gases are collected in the lab, the temperature and the pressure of the gas may change along with a change in the volume.

2. In solving problems of this type, we will always assume that the mass of gas remains constant, that is, the number of moles remains unchanged.

3. The formula that allows for the change of the temperature, pressure and volume is called the Universal Gas Law:

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

Example:

A quantity of gas is found to measure 800 mL when the pressure is 101.3 kPa and the temperature is 20.0°C. The next day the pressure had changed to 125.2 kPa and the temperature was 25.0°C. What was the new volume?

Solution:

Make a list of the variables and their corresponding values:

V_1	= 800 mL	V_2	= unknown
P_1	= 101.3 kPa	P_2	= 125.2 kPa
T_1	= 20.0°C	T_2	= 25.0°C

Remember to change the temperature to the Kelvin scale.

$T_1 = 20.0^o + 273 = 293\ K$ $T_2 = 25.0^o + 273 = 298\ K$

Rearrange the formula to solve for the unknown: $V_2 = \dfrac{P_1V_1T_2}{P_2T_1}$

Substitute the values: $$V_2 = \frac{(800\ mL)(101.3\ kPa)(298\ K)}{(125.2\ kPa)(293\ K)}$$

$V_2 = 658.3\ mL$

2.4	**Determine the relationship between equal volumes of various gases and the molecules they contain.**
2.5	**Express mathematically the results of the analysis of the behaviors of a gas under different conditions.**
2.6	**Determine the ideal-gas constant.**
2.7	**Identify an unknown gas.**
2.9	**Apply the knowledge of the behavior of gases to solve problems.**

KEY CONCEPTS

QUANTITY – VOLUME RELATIONSHIP: AVOGADRO'S LAW

As we add gas to a balloon, we see that the expansion of the balloon depends not only on the pressure and the temperature, but also on the quantity of gas present. Two chemists, Gay-Lussac and Avogadro pioneered this work. The work of Gay-Lussac led Avogadro to propose his famous hypothesis: *Equal volumes of gases at the same temperature and pressure contain equal numbers of molecules*.

Avogadro's Law follows from Avogadro's hypothesis. It states that the volume of a gas maintained at constant temperature and pressure is directly proportional to the number of moles of the gas. If we represent the volume by V and the number of moles by n, then the mathematical relationship between these two values is given by: *Volume = n · constant*. Mathematically this is written: $V = n \cdot k$

From this relationship we can see that doubling the number of moles of gas will cause the volume of the gas to double if temperature and pressure remain constant.

IDEAL GAS EQUATION

Chemists use the expression "ideal gas" to describe a gas whose molecules behave according to the assumptions of the Kinetic Molecular Theory (see page 2-18). Although no real gas completely satisfies this model, most gases can be treated as "ideal" at relatively high temperatures and low pressures (conditions at which they are not able to condense). Under these conditions, Charles' Law, Boyle's Law and Avagadro's Law can be combined to generate a new relationship, the Ideal Gas Equation.

Boyle's law:	$V \propto 1/P$	(constant n, T)
Charles's law:	$V \propto T$	(constant n, P)
Avogadro's law:	$V \propto n$	(constant P, T)

These laws may be combined to produce:

$$V \propto \frac{nT}{P}$$

This expression can also be written as:

$$V = \frac{knT}{P}$$

If we replace the constant of proportionality k, by the letter R (called the Universal Gas Constant), and rearrange this equation mathematically, our relationship becomes the familiar Ideal Gas Equation: PV = nRT.

R has been experimentally determined to equal 8.31 kPa·L/mol·K.

We define Standard Temperature and Pressure (S.T.P.) to be 101.3 kPa and 273 K (0^oC).

If we apply the Ideal Gas Equation for one mole of gas at S.T.P., we have

$$V = nRT/P = (1\text{ mol})(8.3\text{ kPa·L}/\text{mol·K})(273\text{ K})/101.3\text{kPa} = 22.4\text{ L}$$

Therefore, one mole of an ideal gas at S.T.P. occupies a volume of 22.4 L.

At low temperatures and high pressures, real gases do not behave like ideal gases and do not obey these relationships.

Example 1:

If 8.00 g of oxygen gas is found at a pressure of 200 kPa and a temperature of 15.0^oC, how many liters of oxygen are present?

Solution:

V	= unknown
T	= 15.0^oC + 273 = 288 K
P	= 200 kPa
R	= 8.31 kPa·L/mol·K
N	= (8.00g O_2)/32 g/mol = 0.25 mol

Since
PV = nRT,

$$V = \frac{nRT}{P}$$

$$V = \frac{(0.25\text{ mol})\,(8.31\text{ kPa}\cdot\text{L/mol}\cdot\text{K})(288\text{ K})}{200\text{ kPa}}$$

V = 2.99 L

Example 2:

The average size lung of a human body can contain 1.05 L of gas when measured at 37.0^{o}C and a pressure of 98.6 kPa. What is the maximum number of moles of air that are present?

Solution:

V = 1.05 L T = 37.0^{o}C + 273 = 310 K
P = 98.6 kPa R = 8.31 kPa·L/mol·K
PV = nRT
solving for n: n = PV/RT
hence n = (98.6 kPa)(1.05 L)/(8.31 kPa·L/mol·K)(310 K)
n = 0.0402 mol

Example 3:

A student reacts 6.80 g of aluminum carbonate, $Al_2(CO_3)_3$, with an excess of hydrochloric acid, HCl, and collects the carbon dioxide produced. The carbon dioxide collected is under a pressure of 125 kPa and at a temperature of 28.0^{o}C. The balanced equation for this reaction is:

$Al_2(CO_3)_3(s)$ + 6 HCl(aq) $\rightarrow$ 2 $AlCl_3(s)$ + 3 $CO_2(g)$ + 3 $H_2O(l)$
What volume of carbon dioxide gas should this student have collected?

Solution:

$Al_2(CO_3)_3(s)$ + 6 HCl(aq) $\rightarrow$ 2 $AlCl_3(s)$ + 3 $CO_2(g)$ + 3 $H_2O(l)$
Molar mass of aluminum carbonate = 234 g/mol
Given 6.80 g of $Al_2(CO_3)_3$ = 0.02906 mol of $Al_2(CO_3)_3$
From the balanced equation, the mole ratio of aluminum carbonate to carbon dioxide is 1:3, hence there must be 0.08718 mol of CO_2 present.
Using the ideal gas equation, PV = nRT, the volume of the CO_2 may be calculated by V = nRT/P
n = 0.08718 mol R = 8.31 kPa·L/mol·L
T = 28.0^{o}C + 273 = 301 K P = 125 kPa
V = (0.08718 mol)(8.31 kPa·L/mol·K)(301 K)/(125 kPa)
V = 1.74 L of CO_2 collected.

Example 4:

What would be the temperature of 0.20 moles of helium (He) occupying a volume of 64.0 L at a pressure of 50.65 kPa?

Solution:

n = 0.20 mol; V = 64 L; P = 50.65 kPa; R = 8.31 kPa·L/mol·K; T = ?
PV = nRT T = PV/nR
T = (50.65 kPa)(64 L)/(0.20 mol)(8.31 kPa·L/mol·K)
T = 1950. 4 K or T = 1677°C

SAMPLE QUESTIONS

1. *A container holds 43.8 L of chlorine gas, $Cl_2(g)$, at a temperature of 43.0°C and a pressure of 105 kPa.*

   How many moles of gas are present?

   A) 1.75 mol B) 12.9 mol C) 14.6 mol D) 107 mol

2. *A syringe contains 30.0 mL of methane gas, $CH_4(g)$ at a pressure of 105 kPa. The pressure is then reduced to 90.0 kPa while the temperature remains constant.*

   What is the new volume of the methane gas?

   A) 4.00 x 10^{-2} mL C) 3.50 x 10^{1} mL
   B) 2.57 x 10^{1} mL D) 3.15 x 10^{2} mL

3. *The volume of a gas in a balloon changes in each of the following situations:*

   1) The balloon is punctured.
   2) The balloon rises in the atmosphere. (Assume temperature remains constant.)
   3) The balloon is cooled at a constant pressure.
   4) Gas is added to the balloon at a constant temperature and pressure.

   In which situation does the volume of the balloon increase?

   A) 1 and 2 only B) 1 and 3 only C) 2 and 4 only D) 3 and 4 only

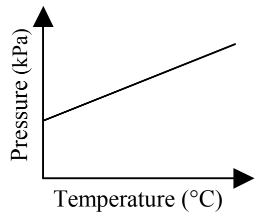
4. *An evacuated 100 L container weighs 480 g. When this container is filled with nitrogen gas, N_2, its total mass is 620 g. When it is filled with an unknown gas at the same temperature and pressure, its total mass is 770 g.*

   Which of the following is the unknown gas?

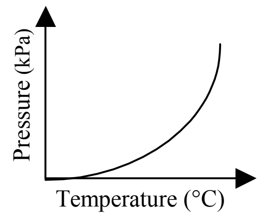
   A) Acetylene, C_2H_2 ***C) Ethane, C_2H_6***
   B) Butane, C_4H_{10} ***D) Methane, CH_4***

5. *Which of the following graphs correctly shows the relationship between the pressure exerted by a sample of gas and its temperature? (Assume the number of moles of gas and the volume remain constants.)*

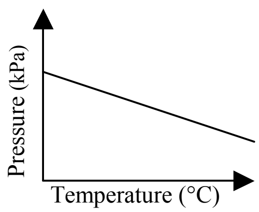
A)



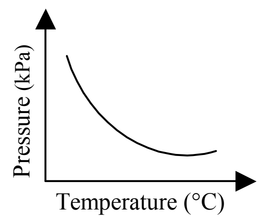
C)



B)

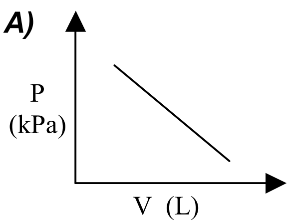


D)

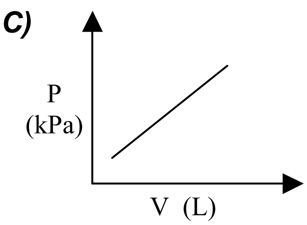


6. *Which one of the following graphs best illustrates the relationship between the volume that a gas occupies and the pressure that the gas exerts on the walls of its container when the number of moles of gas and the temperature remain constant?*

A)



C)



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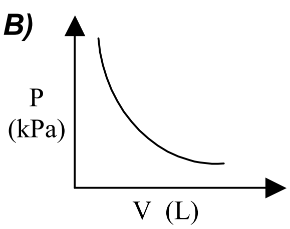
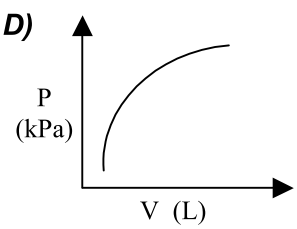


D)



7. *An unmanned space probe lands on Mars. It extends a device that takes a sample of the Martian atmosphere. The sample occupies a volume of 4.00×10^{2} mL at $-123^{o}C$.*

   What volume does the same sample occupy at $27.0^{o}C$ and the same pressure?

   A) 8.00×10^{-1} mL
   B) 8.80×10^{1} mL
   C) 8.00×10^{2} mL
   D) 1.80×10^{3} mL

8. *A sample of helium gas is placed in a 4.0 L container and the pressure is recorded. It is then completely pumped out of the 4.0 L container and into a 0.50 L container, while the temperature is held constant.*

   Which of the following correctly describes what happens to the pressure of the helium gas when it is removed from the 4.0 L container and placed into the 0.50 L container?

   A) The pressure of the helium decreases by a factor of two.
   B) The pressure of the helium decreases by a factor of eight.
   C) The pressure of the helium increases by a factor of two.
   D) The pressure of the helium increases by a factor of eight.

9. *You are to determine the molar mass of an unknown gas. To do this, you collect the following information:*

Volume of the syringe	*113 mL*
Mass of the evacuated syringe	*80.77 g*
Mass of the syringe filled with oxygen gas, O_2	*80.92 g*
Mass of the syringe filled with the unknown gas	*81.07 g*
Temperature	*$22.0^{o}C$*
Pressure	*101.3 kPa*

   According to this information, what is the molar mass of the unknown gas?

10. *A cylinder contains a certain amount of gas at a particular temperature. The volume is reduced to one-half while the Kelvin temperature is tripled. What happens to the pressure of the gas?*

11. *In the laboratory, at a pressure of 100.0 kPa and a temperature of $22.0^{o}C$, you react 30.06 g of calcium, Ca(s), with sufficient hydrochloric acid, HCl(aq), to assure a complete reaction. You obtain calcium chloride, $CaCl_2$(aq) and hydrogen gas, H_2(g). What volume of gas do you obtain?*

12. *Two identical steel cylinders were filled with gases at the same temperature. The first contains 10.0 g of nitrogen gas, N_2 , and the second, 12.0 g of carbon dioxide, CO_2. In which cylinder is the pressure greater?*

13. *A barbecue used for outdoor cooking runs on propane gas, C_3H_8 , contained in a steel cylinder. The cylinder holds 801 g of propane. The equation for the combustion of propane is as follows:*

$$C_3H_8(g) + 5\ O_2\ (g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (g)$$

Given an atmospheric pressure of 101 kPa and an outdoor temperature of 25.0^oC, what volume of oxygen gas, $O_2(g)$ is required to burn all the propane in the cylinder?

14. *Mark is given a sample of gas in the laboratory. He assumes that this gas behaves like an ideal gas. To test his assumption, he conducts an experiment and makes the following observations:*

Number of moles of gas	*2.0 mol*
Volume of gas	*10.0 L*
Temperature	*-73.0^oC*
Pressure	*404 kPa*

Given the above information, is his assumption correct?

15. *A student is given a sample of gas and told that it could be sulfur dioxide (SO_2), nitrogen dioxide (NO_2), or dinitrogen pentoxide (N_2O_5). To determine the identity of the gas, the student evacuates a cylinder of unknown volume and finds its mass. He then fills the cylinder with oxygen gas (O_2) and finds its mass. Finally, he evacuates the cylinder, refills it with the unknown gas under the same conditions as when oxygen was used and finds its mass. Here is his data:*

Mass of evacuated cylinder	*= 143.25 g*
Mass of cylinder and oxygen gas	*= 148.78 g*
Mass of cylinder and unknown gas	*= 161.91 g*

Identify the unknown gas.

16. *A student reacts an unknown mass of iron (Fe) in an excess of sulfuric acid (H_2SO_4). The balanced equation is:*

$$2\ Fe(s) + 3sH_2SO_4(aq) \rightarrow Fe_2(SO_4)_3(aq) + 3\ H_2(g)$$

She records the following data:

Temperature	*23.0^oC*
Pressure	*98.4 kPa*
Volume of dry hydrogen collected	*47.3 mL*

Find the mass of iron used in the experiment.

17. *A cylinder with a movable piston contains 8.40 g of neon gas (Ne) at a temperature of 22.0^{o}C. A second identical cylinder contains 8.40 g of carbon dioxide gas (CO_2). The cylinder containing carbon dioxide is heated while keeping its pressure the same as the pressure in the cylinder containing the neon. To what temperature (in ^{o}C) will the carbon dioxide cylinder have to be heated so that it will occupy the same volume as the cylinder containing the neon?*

18. *The following equation illustrates the reaction that takes place when automobile air bags are activated:*

    $$2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$$

    Nitrogen gas (N_2) is produced to fill the air bags and provides the required cushioning effect. An air bag holds a maximum of 5.60 x 10^4 mL of nitrogen at a temperature of 25.0^{o}C and a pressure of 101.3 kPa. What mass of sodium azide ($NaN_3(s)$) will have to be placed in the air bag to produce 5.60 x 10^4 mL of nitrogen at these conditions?

19. *An environmental biologist takes an air sample outside an oil refining plant on a hot summer day. The sample occupies a volume of 1.50 x 10^3 mL at 31.0^{o}C and 101.8 kPa. The sample is then taken indoors where it is drawn off into a large empty container with a volume of 2.00 L. The gas has a pressure of 72.3 kPa. What change in temperature has the gas undergone?*

20. *A limnologist is a biologist who studies water ecosystems such as marshes, bogs and lakes. A limnologist studying the ecosystem in a marsh notices bubbles coming up through the marsh water. She traps the gas in a syringe and returns to the laboratory. Below is the data that was recorded in the lab:*

Volume of the syringe	*175.0 mL*
Mass of the evacuated syringe	*50.78 g*
Mass of syringe and unknown gas	*50.89 g*
Temperature of the gas	*21.0^{o}C*
Pressure of the gas	*96.0 kPa*

From this information, determine which of the following gases could be the one trapped by the limnologist in the syringe.

1) *oxygen gas (O_2)*
2) *nitrogen gas (N_2)*
3) *sulfur dioxide gas (SO_2)*
4) *methane gas (CH_4)*
5) *carbon dioxide gas (CO_2)*

21. *The discovery of oxygen gas is often accredited to an English clergyman by the name of Joseph Priestley. In 1774 he reported its discovery and described some of its properties. He discovered oxygen by heating a red powder called "calx of mercury". This substance is now known as mercury (II) oxide (HgO). The equation for this reaction is:* $2\ HgO(s) \rightarrow 2\ Hg(l) + O_2(g)$.

    What volume of oxygen gas, measured at 30.0°C and 95.0 kPa can be produced from the complete decomposition of 8.66 grams of mercury (II) oxide?

22. *During the winter, a driver inflates his car tire to a pressure of 325 kPa when the outdoor temperature is – 12.0°C. On a trip during the summer, heat from the road and friction cause the air in the tire to reach a temperature of 45.0°C. Assume that none of the air leaked out of the tire during the change in seasons and the volume of the tire does not change. What is the pressure of the tire during the summer trip?*

3.1 Identify the uses of substances in a given phase, in natural or man made objects, based on the characteristic properties of the phases of matter.
3.2 Interpret the behavior of matter in each of its phases, based on the molecular model and the knowledge of gases.
3.3 Illustrate possible molecular motions in the different phases of matter.

KEY CONCEPTS

KINETIC MOLECULAR THEORY

The ideal gas equation, $PV = nRT$, describes the relationship between volume, temperature, pressure and the number of moles of gas present, but it does not tell us how gases behave. To understand this idea, we need the *Kinetic Molecular Theory*, which describes the movement of molecules. This theory states:

1. All gases are made up of large numbers of molecules, which are in constant random motion.

2. The volume of all of the molecules of the gas is negligible when compared to the total volume in which the gas is contained.

3. Molecules can collide with each other with no loss of kinetic energy. In other words, the collisions are perfectly elastic.

4. The forces of attraction and repulsion between the actual gas molecules are negligible.

5. The average kinetic energy of the molecules is proportional to the absolute temperature. At any given temperature, the molecules of all gases have the same average kinetic energy.

According to the kinetic molecular theory, the pressure that is exerted by a gas on the walls of its container is the result of the continual hitting of the molecules against the container wall. Every time a molecule hits the container wall, it causes a force to be exerted against the wall. Pressure is defined as force per unit area and hence it follows that the pressure present in the container is a direct result of these collisions on a given area.

We can see that if we were to halve the volume of a particular container filled with some gas, the gas molecules would make twice as many collisions per unit area and the result would be a doubling of the pressure. The kinetic theory provides us with a simple explanation of Boyle's Law that pressure and volume are inversely related.

The kinetic theory also explains why the volume of a gas increases with increases in the temperature. If a gas is contained in a cylinder with a movable piston, increasing the temperature of the gas causes the average velocity of the gas molecules to increase. This will cause an increase in the bombardment of the molecules against the piston. Since the molecules will have greater energy, they will exert a greater force on the piston causing it to move outward, away from the bombardment of the molecules and hence increasing the volume of the container. The kinetic molecular theory thus provides us with an explanation of Charles' Law.

Example:

Place the following gases in order of decreasing average molecular speed at 25.0^{o}C:
CO, Cl_2, HI, SF_6, H_2S

> ***Solution:***
> *CO > H_2S > Cl_2 > HI > SF_6. The solution is based on the idea that heavier molecules move more slowly. Calculating the molar mass of each gas indicates that CO has the smallest molar mass of 28 g/mol and SF_6 has the greatest molar mass of 146 g/mol. If we calculate the molar masses of the other given gases, we can rank them in increasing order.*

SAMPLE QUESTIONS

1. *Which of the following statements is true?*

   A) At the same temperature and pressure, two samples of gases with different molar masses will have different average molecular velocities.
   B) Temperature has no effect on the velocity of gas molecules.
   C) Molecules move more slowly in the gaseous phase than in the aqueous phase.
   D) At the same temperature, molecules of different gases move at the same velocity.

2. *Which of the following is NOT part of the kinetic molecular theory relating to ideal gases.*

   A) Gases are made of tiny particles called molecules.
   B) Gas molecules in an ideal gas show a great attraction for each other and are constantly undergoing inelastic collisions.
   C) Gas molecules in an ideal gas have no volume.
   D) Pressure results from the collision of the gas molecules with the walls of the container.

3. *A 2.0 L container holds pure hydrogen at 25.0^{o}C and 101 kPa of pressure. A second 2.0 L container holds pure oxygen gas at 25.0^{o}C and 101 kPa of pressure.*

   Which of the following statements concerning these two samples of gas are true?

   1) *The average velocity of the oxygen molecules is equal to the average velocity of the hydrogen molecules.*
   2) *The average kinetic energy of the oxygen molecules is equal to the average kinetic energy of the hydrogen molecules.*
   3) *The mass of oxygen present is equal to the mass of hydrogen present.*
   4) *The number of oxygen molecules is equal to the number of hydrogen molecules.*

   A) 1 and 2 B) 1 and 3 C) 2 and 3 D) 2 and 4

MODULE 3
- ENERGY IN CHEMICAL REACTIONS -

Review from Physical Science 416/436

Physical change: A physical change does not change the composition of the substance. Only the appearance or state of the substance is changed. Usually, physical changes can be reversed by physical means.

Examples:
A **change of state** (solid/liquid/gas) can be reversed by adding or removing energy: the substance is still the same substance.

A **liquid-liquid solution** (homogeneous mixture of two liquids) may be separated into its constituent substances by a physical process such as distillation. Making the solution did not change the composition of each of the constituent substances.

Chemical change: A chemical change involves a **change in composition** of the substance into another substance. Chemical changes cannot be reversed by physical means. If reversible, they can be reversed only by chemical means.

Example:
Any substance which reacts with oxygen undergoes a chemical change. This includes burning, rusting and tarnishing.

<u>Laboratory observations which may indicate a chemical change has occurred</u>:

- Effervescence (bubbles)
- Precipitate formation
- Colour change
- Heat released or absorbed
- Light given off or absorbed
- Electricity produced

Nuclear change: A nuclear change involves a change in the nuclear structure of the substance. Nuclear changes cannot be reversed.

Example:
Formation of an isotope of an element (addition or removal of neutrons from the nucleus).

Each type of change above involves a transfer of energy, either to or from the surrounding environment. The following are listed in increasing order as to the amount of energy transferred: **physical change**, **chemical change** and **nuclear change**.

SAMPLE QUESTIONS

1. *Which of the following are chemical changes?*

   1) *The production of water vapour and carbon dioxide as a result of burning wood in a wood stove.*
   2) *The melting of butter in a microwave oven.*
   3) *The explosion of dynamite in an iron ore mine.*
   4) *The sublimation of moth balls in a cedar chest.*
   5) *The formation of raindrops in a cloud.*
   6) *The disintegration of uranium-235 in a nuclear reactor.*
   7) *The electrolysis of water.*

   A) 1, 2 and 6 ***C) 1, 3 and 5***
   B) 1, 3 and 7 ***D) 2, 4 and 7***

2. *The reactions which typically involve the largest quantities of heat are:*

   A) thermochemical ***C) changes of state***
   B) nuclear ***D) reversible reactions***

3.1 Associate the enthalpy of a substance with the kinetic and potential energy of its molecules.

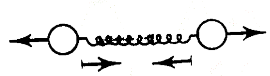
KEY CONCEPTS

Energy:	Energy is the capacity to do work.
Law of Conservation of Energy	In any chemical or physical process, energy is neither created nor destroyed. It is converted from one form of energy to another.

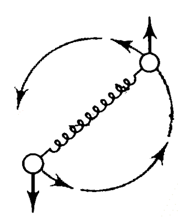
Forms of Energy:

Potential Energy (E_p):	stored energy due to position.
Kinetic Energy (E_k):	the energy of motion.
Heat (Thermal) Energy (Q):	the energy which is transferred from one body to another due to a difference in temperature between the two bodies.
Others:	radiant, mechanical, chemical and electrical

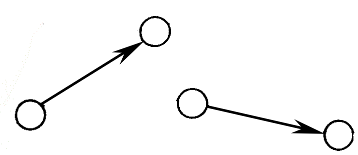
Kinetic Energy of Particles:
Particles (atoms, molecules or formula units) exhibit three (3) types of kinetic energy, depending on their states of matter:



vibrational E_k



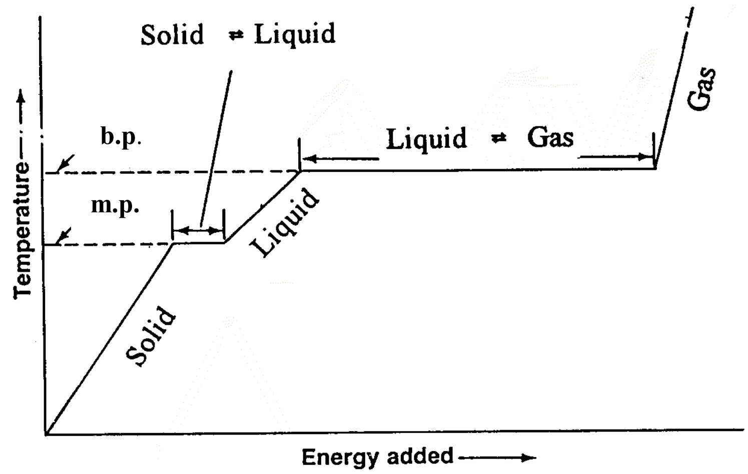
rotational E_k



translational E_k

Solids mainly exhibit:	vibrational E_k
Liquids mainly exhibit:	vibrational E_k and rotational E_k
Gases mainly exhibit:	vibrational E_k, rotational E_k and translational E_k

Phase Diagram:



Note that at both the melting point (m.p.) and boiling point (b.p.), a certain amount of energy can be added to the system without a corresponding increase in temperature.

SAMPLE QUESTION

1. *The type of energy that a molecule exhibits where the atoms of the molecule move alternately toward and away from the molecule's centre of mass is called:*

   A) Rotational kinetic energy
   B) Translational kinetic energy
   C) Vibrational kinetic energy
   D) Potential kinetic energy

3.2	**Associate the heat of reaction of a chemical reaction with changes in the enthalpy of the reactants and the enthalpy of the products.**
3.3	**Illustrate, using graphs, the enthalpy change of substances in an endothermic chemical reaction and in an exothermic chemical reaction after observing demonstrations.**
1.1	**Identify chemical and physical changes that release more energy than they absorb, based on what they have observed in their environment and in the Laboratory.**
1.2	**Identify, based on what they have observed in their environment and in the laboratory experiments, chemical and physical changes that absorb more energy than they release.**
1.3	**Classify chemical and physical changes as either exothermic or endothermic, after observing demonstrations and carrying out experiments.**

KEY CONCEPTS

Enthalpy (H): the heat content of a substance, or the energy stored in a substance during its formation. [Since H cannot be measured directly, we usually refer to the Change in Enthalpy, ΔH]

Heat of Reaction (ΔH): the amount of energy absorbed or released during a reaction.

Heat of Formation (ΔH_f): the amount of energy absorbed or released when a compound is formed from its elements.

Heat of Solution (ΔH): the amount of energy absorbed or released when a solute dissolves completely in a solvent.

Heat of Combustion (ΔH): the amount of energy released when a substance burns completely.

Other "Heats" include the Heat of Fusion, Heat of Vaporization and Heat of Neutralization. "Molar" Heats refer to the amount of energy absorbed or released when one (1) mole of the substance under consideration undergoes a change or is produced.

Physical or chemical reactions involve a net intake (absorption) or release of energy. The net change in enthalpy, ΔH, is the difference between the Heat of Formation (ΔH_f) of the **Products** in the reaction and the Heat of Formation (ΔH_f) of the **Reactants** in the reaction:

$$\Delta H \quad = \quad \Delta H_f \text{(Products)} \quad - \quad \Delta H_f \text{(Reactants)}$$

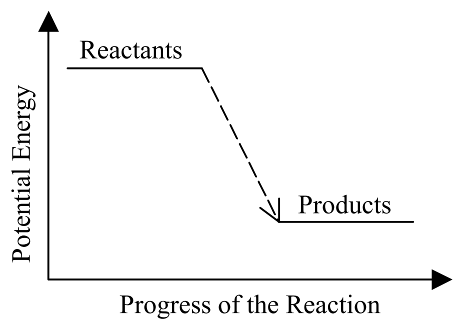
If the ΔH of a reaction is **negative**, then the reaction is **exothermic**, and heat is **released**.

Exothermic reactions in everyday life are those which involve release of the energy, such as the combustion of gasoline, or human metabolism.

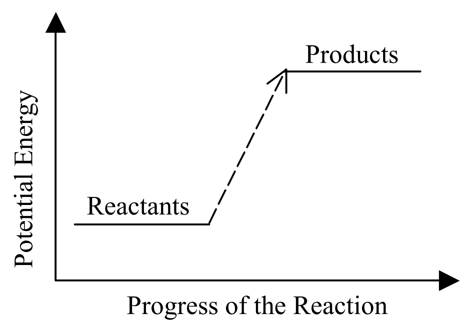
If the ΔH of a reaction is **positive**, then the reaction is **endothermic**, and heat is **absorbed**.

Endothermic reactions in everyday life are those which involve the process of storing energy, such as charging a battery, or photosynthesis.

Potential Energy (Enthalpy) Diagrams for Exothermic and Endothermic Reactions:



EXOTHERMIC
$\Delta H < 0$



ENDOTHERMIC
$\Delta H > 0$

SAMPLE QUESTIONS

1. *Which of the following statements referring to enthalpy is true?*

   A) Enthalpy always decreases during a chemical reaction.
   B) Enthalpy always increases during a change of state.
   C) Enthalpy always remains unchanged during the formation of chemical bonds.
   D) Enthalpy always decreases during an exothermic change.

2. *The molar heat of formation of a compound is*

   A) the heat required to atomize one mole of that compound
   B) the heat liberated during the bond formation of one mole of that compound
   C) the heat of reaction for the process of making one mole of that compound from the reaction between two other compounds.
   D) the heat of reaction for the process of making one mole of that compound from its elements.

3. *For the following, which process releases more energy than it absorbs?*

   A) Water cools down when $NH_4Cl_{(s)}$ is dissolved in it.
   B) Solid margarine melts on a hot plate.
   C) A car's engine burns gasoline.
   D) River water evaporates.

4. *For which of the following situations is more energy absorbed than released?*

   1) Natural gas (CH_4) is burned in a furnace.
   2) When solid KBr is dissolved in water, the solution gets colder
   3) Water is boiled in a tea kettle
   4) When concentrated sulfuric acid (H_2SO_4) is added to water, the solution gets very hot.
   5) When a test tube containing iodine crystals is lowered in hot water, purple vapour forms.

   A) 1, 2 and 3 ***C) 2, 4 and 5***
   B) 2, 3 and 5 ***D) 1, 3 and 5***

5. *Which of the following involves an exothermic change?*

   A) Ice cream melting ***C) Defrosting food***
   B) Wet leaves drying ***D) Frost forming on windows***

6. *The ΔH of a reaction is negative if:*

   1) the reaction is endothermic.
   2) the reaction is exothermic.
   3) the enthalpy of the products is greater than that of the reactants.
   4) the enthalpy of the products is less than that of the reactants.

   A) 2 and 4 ***B) 2 and 3*** ***C) 1 and 4*** ***D) 1 and 3***

7. *The following equation represents a reaction:*

   $A + B \rightarrow C + x\ kJ$

   Which of the following statements is true?

   A) The enthalpy of the product is less than the enthalpy of the reactants.
   B) The enthalpy of the product may be greater than or equal to the enthalpy of the reactants.
   C) The enthalpy of the product may be less than or equal to the enthalpy of the reactants.
   D) The enthalpy of the product is greater than the enthalpy of the reactants.

8. *Based on observations which you made in the laboratory, which of the following statements are true?*

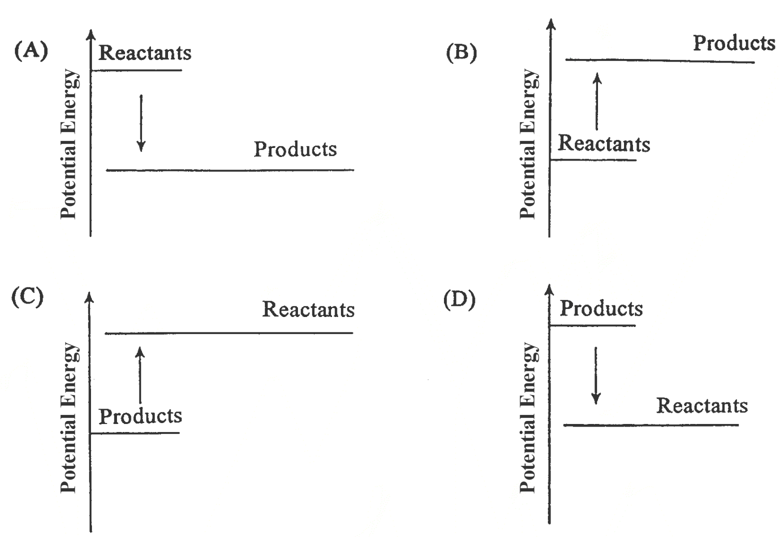
   1) *Energy is conserved during a transformation of energy from one form to another.*
   2) *In an energy diagram, energy released during a chemical reaction is indicated with a negative value.*
   3) *Energy is not conserved during a transformation of energy from one form to another.*
   4) *We can determine both the equation and the overall energy of a given thermochemical equation by adding partial thermochemical equations for the given reaction.*
   5) *In an energy diagram, energy absorbed during a chemical reaction is indicated with a positive value.*

   A) 1, 2, 3 and 5 ***C) 1, 2, 4 and 5***
   B) 2, 3, 4 and 5 ***D) 1, 2, 3 and 4***

9. *Which of the following reactions is endothermic?*

   A) $N_2 + 3H_2 \rightarrow 2NH_3$ $\Delta H = -92\ kJ$
   B) $2H_2 + O_2 \rightarrow 2H_2O + heat$
   C) $C + O_2 \rightarrow CO_2 + 394\ kJ$
   D) $N_2 + 2O_2 + 67.6\ kJ \rightarrow 2NO_2$

10. *Amongst the following energy diagrams, which one correctly represents an endothermic reaction?*



11. *When solid ammonium chloride (NH_4Cl) is added to water, it forms an aqueous solution which feels cold to the touch. Which one of the following best accounts for this observation?*

   A) Heat is released from the system so it feels colder.
   B) $NH_4Cl_{(s)} \rightarrow NH_4Cl_{(aq)}$ $\Delta H = +32.1$ kJ
   C) The reaction is exothermic.
   D) $NH_4Cl_{(s)} \rightarrow NH_4Cl_{(aq)} + 32.1$ kJ

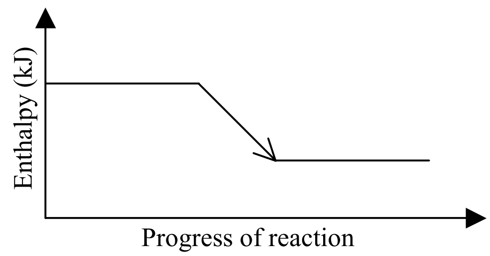
12. *A student dissolves some KOH(s) in a flask of water, which is represented by the following equation:*

$$KOH_{(s)} \rightarrow K^{+}\,(aq) + OH^{-}\,(aq)$$

   The temperature of the liquid in the flask increases from 14.8°C to 24.3°C. Which of the following statements describes this situation?

   A) The enthalpy change for the process: $KOH_{(s)} \rightarrow K^{+}\,(aq) + OH^{-}\,(aq)$ has a positive value.
   B) The potential energy (enthalpy) of $KOH_{(s)}$ is higher than the potential energy (enthalpy) of $K^{+}\,(aq) + OH^{-}\,(aq)$.
   C) The dissolving of $KOH_{(s)}$ in water is an endothermic process.
   D) Heat energy is transferred from the surroundings to the system as the NaOH(s) dissolves.

13. The diagram below illustrates the enthalpy diagram for a chemical reaction: Which of the following two equations can be represented by this diagram?



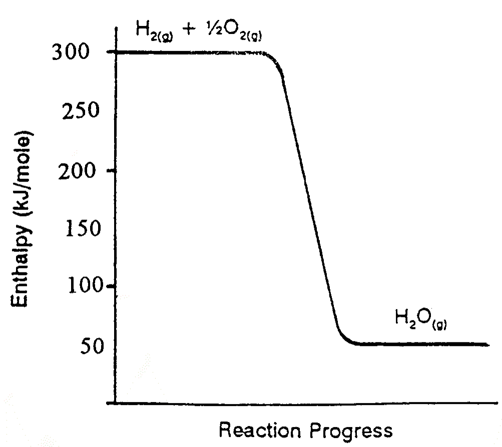
1) $C_{(s)} + H_2O_{(g)} \rightarrow CO_{(g)} + H_{2(g)}$ $\Delta H = +132\ kJ$
2) $2\ HCl_{(g)} + 185\ kJ \rightarrow H_{2(g)} + Cl_{2(g)}$
3) $H_2O_{(l)} \rightarrow H_2O_{(s)} + 6.02\ kJ$
4) $N_{2(g)} + 2\ O_{2(g)} \rightarrow 2NO_{2(g)}$ $\Delta H = -68.0\ kJ$

A) 1 and 2
B) 1 and 3
C) 2 and 4
D) 3 and 4

14. Refer to the graph below:

a) Determine the ΔH value for the following reaction:

b) Is this reaction exothermic or endothermic?



2.1	**Analyze the transfer of heat energy that occurs during the formation of a mixture, using measurements and calculations.**
2.6	**Apply their knowledge of the transfer of heat energy to solve problems and do numerical and graphical exercises.**

KEY CONCEPTS

Calorimetry

Calorimeters are instruments used to experimentally determine the heat energy (ΔH) absorbed or released during a given reaction. The reaction is carried out inside the calorimeter in a sealed compartment. Since the calorimeter is a closed system, and the sealed compartment is surrounded by a second compartment containing water, the ΔH can be calculated by measuring the water temperature before and after the reaction.

Specific Heat: the amount of energy required to raise the temperature of one gram of a substance by one degree Celsius.

Example:
Consider the following data recorded for the partial combustion of wax ($C_{25}H_{52}$) in a calorimeter. Using this data, calculate the molar heat of combustion of wax.

Data:	initial mass of wax	22.35 g	
	final mass of wax	12.08 g	
	volume of water in calorimeter	352.5 mL	(Note: 1 mL of water has a mass of 1 g)
	initial temperature of the water	12.6 ^oC	
	final temperature of the water	43.5 ^oC	

Solution:

Step 1: Calculate the amount of energy required to increase the temperature of the water:

$Q = mc\Delta T$, where
q is the quantity of heat which increased the temperature of the water;
m is the mass of the water;
c is the specific heat of water, which is 4.19 J/g^{o}C;
ΔT is the $T_{final} - T_{initial}$ of the water.

$Q = (352.5\text{g})(4.19\ \text{J/g}^o\text{C})(30.9^o\text{C})$
$Q = 4.56 \times 10^4$ J or 45.6 kJ.

Step 2: Calculate the Molar Heat of Combustion of wax.

$$\frac{Q\ (\text{water})}{\Delta m\ (\text{wax})} = \frac{\Delta H\ (\text{wax})}{\text{molar mass (wax)}}$$

$$\frac{45.6\ \text{kJ}}{10.27\ \text{g}} = \frac{\Delta H(\text{wax})}{352.74\ \text{g/mol}}$$

$$\Delta H\ (\text{wax}) = 1.57 \times 10^{3}\ \text{kJ/mol}$$

Heat Transfer

When two bodies of different temperature come into contact, heat is transferred from the body or substance with the higher temperature, to that with the lower temperature until equilibrium is reached, or both are the same temperature. At equilibrium, the quantity of heat lost (Q loss) of the substance which was at the higher temperature must equal the quantity of heat gained (Q gain) of the substance which was at the lower temperature.

Example:
A beaker contains 400.0 mL of water at 60.0°C. A student adds 250.0 mL of water which is at a temperature of 20.0°C. Find the temperature of the final mixture.

Solution

$$-(Q_{loss}) = Q_{gain}$$

$$-Q(60^{o}C_{water}) = Q(20^{o}C_{water})$$

$$-mc\Delta T = mc\Delta T$$

$$-mc(T_f - T_i) = mc(T_f - T_i)$$

$$-(400.0\ g)(4.19\ J/g^{o}C)(T_f - 60.0^{o}C) = (250.0g)(4.19\ J/g^{o}C)(T_f - 20.0^{o}C)$$

$$-(T_f - 60.0) = \frac{250.0(T_f - 20.0)}{400.0}$$

$$-(T_f - 60.0) = 0.625(T_f - 20.0)$$

$$-(T_f - 60.0) = 0.625\ T_f - 12.5$$

$$-T_f - 0.625\ T_f = -12.5 - 60.0$$

$$-1.625\ T_f = -72.5$$

$$T_f = 44.6^{o}\ C$$

SAMPLE QUESTIONS

1. *A sample of aluminum is heated from 26.0^{O}C to 66.0^{O}C in a calorimeter. The energy absorbed by the aluminum is found to be 36.0 kJ. The specific heat capacity of aluminum is 0.900 J/g^{O}C. What mass of aluminum absorbed this heat?*

   A) 1.00×10^{-3} g
   B) 1.00 g
   C) 1.00×10^{1} g
   D) 1.00×10^{3} g

2. *Calculate the molar heat released by dissolving 4.0 g of potassium hydroxide, KOH(s) in 200.0 mL of water, if the water inside the calorimeter was heated from 25.0^{O}C to 31.5^{O}C as a result of the dissolving. (Note: you may assume that both solutions behave like water, and that 1.0 mL of each solution has a mass of 1.0 gram.)*

   A) 4.4 kJ/mol
   B) 14 kJ/mol
   C) 1.4 kJ/mol
   D) 76 kJ/mol

3. *When 200.0 mL of 0.50 mol/L NaOH solution completely neutralizes 200.0 mL of 0.50 mol/L HCl solution, the temperature rises from 20.8^{O}C to 27.4^{O}C. Calculate the molar heat of neutralization of NaOH. (Note: you may assume that both solutions behave like water, and that 1.0 mL of each solution has a mass of 1.0 gram.)*

   A) 1.1×10^{2} kJ/mol
   B) 2.2×10^{2} kJ/mol
   C) 1.1×10^{3} kJ/mol
   D) 5.5×10^{3} kJ/mol

4. *Determine the final temperature of water, after the combustion of $CH_{4(g)}$ in a bomb calorimeter, given the following data:*

   Mass of methane burned: 1.00 g
   Volume of water in the bomb calorimeter: 800.0 mL
   Heat released from the combustion of the gas: 50.4 kJ
   Initial temperature of the water: 27.0^{O}C

   A) 15.3^{O}C
   B) 25.7^{O}C
   C) 42.1^{O}C
   D) 52.5^{O}C

5. *Calculate the heat released by the combustion of one mole of butane, $C_4H_{10}(g)$, based on the following experimental data:*

initial mass of butane: 7.25 g — *initial temperature of the water: 25.0^{O}C*
final mass of butane: 1.25 g — *final temperature of the water: 60.7^{O}C*
volume of water in calorimeter: 2.00 L

A) 298 kJ/mol ***C) 50.0 kJ/mol***
B) 289 kJ/mol ***D) 2890 kJ/mol***

6. *A student mixes a sample of Liquid A with a sample of Liquid B. Assume no heat is lost to the surroundings. Which of the following factors affect the final temperature of this mixture?*

   1) *The masses of each liquid used.*
   2) *The molar mass of each liquid.*
   3) *The specific heat capacity of each liquid.*
   4) *The initial temperatures of each liquid.*
   5) *The atmospheric pressure of the environment.*

A) 1, 2 and 4 only ***C) 2, 3 and 4 only***
B) 1, 3 and 4 only ***D) 3, 4 and 5 only***

7. *If a mixture is made consisting of 100.0 mL of water at 90.0^{O}C and 100.0 mL of water at 25.0^{O}C, what will the final temperature of the mixture be? (Assume no heat loss to the surroundings.)*

A) 57.5^{O}C ***B) 37.0^{O}C*** ***C) 52.5^{O}C*** ***D) 50.0^{O}C***

8. *How much heat does a 24.6 g block of copper absorb if its temperature increases from 0^{O}C to 25^{O}C? (c_{copper} = 0.39 J/g^{O}C).*

9. *A calorimeter contains 230.0 g of water at 25^{O}C.*
A 200.0 g sample of copper at 47 ^{O}C is placed into the calorimeter.
Assuming complete heat transfer between the copper and the water, calculate the final temperature of the water in the calorimeter.
(Note: c_{copper} = 0.39 J/g^{O}C).

10. *A mining company does regular testing of the quality of coal it mines. Each test consists of burning a 0.600 g sample of coal, C(s), in a calorimeter containing 200.00 mL of water at a temperature of 22.2^{o}C. According to the company, coal of good quality should have a minimum molar heat of combustion of -391 kJ/mol.*

    What minimum temperature should the water in the calorimeter reach if the sample of coal tested is of good quality?

11. *A beaker contains 350.0 mL of water at 60.0^{o}C. A student adds 275.0 mL of water which is at 22.0^{o}C, and then adds another 300.0 mL of water at 70.0^{o}C.*

    a) Find the temperature of the mixture after the first 275.0ml was added.
    b) Find the temperature of the final mixture.

12. *You are to heat up 300.00 mL of strong coffee, which has a temperature of 22^{o}C. To do this, you add 125.00 mL of water, the temperature of which is 98^{o}C. What will the temperature of the coffee be after you do this?*

2.5 Determine the heat of formation of a substance, based on a series of chemical reactions whose heat of reaction were determined by means of an experiment or found in a manual.

KEY CONCEPTS

Hess' Law Of Constant Heat Summation

s

When a chemical reaction can be expressed as the sum of two or more reactions, its heat of reaction (ΔH) is the sum of the heats of reaction of each of the individual reactions.

Example:

We can determine the heat of reaction for the following reaction,

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

given the following thermochemical equations:

1)	$H_2O_{(g)} + C_{(s)} \rightarrow CO_{(g)} + H_{2\,(g)}$	$\Delta H = 131.4$ kJ
2)	$CO_{(g)} + 1/2\ O_{2\,(g)} \rightarrow CO_{2\,(g)}$	$\Delta H = 282.9$ kJ
3)	$H_2O_{(g)} \rightarrow H_{2\,(g)} + 1/2\ O_{2\,(g)}$	$\Delta H = 241.8$ kJ

Solution

Step 1: Rewrite the thermochemical equations so that the molecules which appear in the net equation appear on the same side of the equation as they do in the net equation. If an equation is reversed, the sign of the ΔH must be changed.

1)	$H_2O_{(g)} + C_{(s)} \rightarrow CO_{(g)} + H_{2\,(g)}$	$\Delta H = 131.4$ kJ
2)	$CO_{(g)} + 1/2\ O_{2\,(g)} \rightarrow CO_{2\,(g)}$	$\Delta H = 282.9$ kJ
3)	$H_{2\,(g)} + 1/2\ O_{2\,(g)} \rightarrow H_2O_{(g)}$	$\Delta H = 241.8$ kJ

Step 2: Add the equations together, including the H values, and eliminate or "cancel" molecules which appear both on the right and on the left of <u>any</u> chemical equation arrow ($\rightarrow$).

1)	$H_2O_{(g)}$	+	$C_{(s)}$	$\rightarrow$	$CO_{(g)} + H_{2\,(g)}$	$\Delta H = 131.4$ kJ
2)	$CO_{(g)}$	+	$1/2\ O_{2\,(g)}$	$\rightarrow$	$CO_{2\,(g)}$	$\Delta H = 282.9$ kJ
3)	$H_{2\,(g)}$	+	$1/2\ O_{2\,(g)}$	$\rightarrow$	$H_2O_{(g)}$	$\Delta H = 241.8$ kJ
Add:	$C_{(s)}$	+	$O_{2\,(g)}$	$\rightarrow$	$CO_{2\,(g)}$	$\Delta H = 393.3$ kJ

Note: In order for all molecules which are not required to be "cancelled", it is sometimes necessary to multiply an entire thermochemical equation, including the H, by a factor. Also, it may be necessary to divide the net ΔH by a factor, if you have been asked to calculate a Molar ΔH, and the coefficient for the substance concerned is not one in the net equation.

SAMPLE QUESTIONS

1. *Given the following thermochemical equations:*

   1) $K_{(s)} + 1/2\ Cl_{2(l)} \rightarrow KCl_{(s)}$ $\quad \Delta H = -420$ kJ
   2) $K_{(s)} + 1/2\ Cl_{2(g)} \rightarrow KCl_{(s)}$ $\quad \Delta H = -440$ kJ

   What does the difference between the two heats of reaction, 20 kJ, represent?

   A) The heat required to melt 1 mole of $KCl_{(s)}$
   B) The heat released when 2 moles of $KCl_{(s)}$ ***are formed***
   C) The heat required to condense 0.5 moles of $Cl_{2(l)}$
   D) The heat required to vaporize 0.5 moles of $Cl_{2(l)}$

2. *Given the following equations:*

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} + 394\ kJ$

$H_{2(g)} + 1/2\ O_{2(g)} \rightarrow H_2O_{(g)} + 242\ kJ$

$C_{(s)} + 2\ H_{2(g)} + 1/2\ O_{2(g)} \rightarrow CH_3OH_{(l)} + 239\ kJ$

What amount of heat is released by the combustion of one mole of methanol, $CH_3OH_{(l)}$?

$CH_3OH_{(l)} + 3/2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)}$

A) 639 kJ/mol
B) 1117 kJ/mol
C) 875 kJ/mol
D) 87 kJ/mol

3. *The chemical equation for the formation of carbon disulfide, CS_2, from C and S is as follows:*

$$C + 2\ S \rightarrow CS_2$$

Calculate the ΔH of formation for the above reaction, given the following equations:

$C + O_2 \rightarrow CO_2 + 393.5\ kJ$

$S + O_2 \rightarrow SO_2 + 296.1\ kJ$

$CS_2 + 3O_2 \rightarrow CO_2 + 2SO_2 + 1072\ kJ$

A) 86.3 kJ/mol ***C) 382.4 kJ/mol***
B) -1761.6 kJ/mol ***D) - 2057.7 kJ/mol***

4. *What is the ΔH for the following reaction?*

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

We know the following information:

a) $1/2\ O_{2(g)} + 1/2\ N_{2(g)} \rightarrow NO_{(g)}$ $\Delta H = 90.3\ kJ$
b) $1/2\ N_{2(g)} + O_{2(g)} \rightarrow NO_{2(g)}$ $\Delta H = 33.8\ kJ$

5. *Given that,*

a)	$2\ PCl_{3(l)} + O_{2(g)} \rightarrow 2\ POCl_{3(l)}$	$\Delta H = -587$ kJ/mol
b)	$P_4O_{10(s)} + 6\ PCl_{5(s)} \rightarrow 10\ POCl_{3(l)}$	$\Delta H = -419$ kJ/mol
c)	$2\ P_{(s)} + 3\ Cl_{2(g)} \rightarrow 2\ PCl_{3(l)}$	$\Delta H = -686$ kJ/mol
d)	$2\ P_{(s)} + 5\ Cl_{2(g)} \rightarrow 2\ PCl_{5(s)}$	$\Delta H = -892$ kJ/mol

Calculate the enthalpy of formation of tetraphosphorous decoxide,

$$4\ P_{(s)} + 5\ O_{2(g)} \rightarrow P_4O_{10(s)}$$

6. *The bombardier beetle uses an explosive discharge as a defensive measure. The chemical reaction responsible for the defensive discharge is the oxidation of hydroquinone by hydrogen peroxide to produce quinone and water.*

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

Calculate the enthalpy change, ΔH, for this reaction from the following data:

$C_6H_4(OH)_{2(aq)} \rightarrow C_6H_4O_{2(aq)} + H_{2(g)}$	$\Delta H = +177.4$ kJ
$H_{2(g)} + O_{2(g)} \rightarrow H_2O_{2(aq)}$	$\Delta H = -191.2$ kJ
$H_{2(g)} + 1/2\ O_{2(g)} \rightarrow H_2O_{(g)}$	$\Delta H = -242.8$ kJ
$H_2O_{(g)} \rightarrow H_2O_{(l)}$	$\Delta H = -43.8$ kJ

MODULE 4

- RATE OF CHEMICAL REACTIONS -

1.1 Compare the changes that substances in the environment undergo, according to the rate at which these changes occur.

KEY CONCEPTS

A **chemical change** produces a new kind of matter (products) with different properties than the original substances (reactants).

Chemical reactions occur at different rates.

examples:

slow reaction:	a copper roof changes from a copper colour to green over a period of several years
fast reaction:	a magnesium strip burns with a bright white light if heated by a flame
explosive reaction:	a mixture of hydrogen and oxygen explodes when exposed to a flame

Note: The classification of slow, average, fast and explosive is relative only and does not relate to a specific reaction rate.

SAMPLE QUESTION

1. Arrange the following chemical reactions from fastest to slowest.

1) a piece of sodium reacts in a beaker of distilled water
2) a man fires a shot gun
3) a compost pile decomposes

A) 1, 3, 2
B) 2, 1, 3
C) 3, 1, 2
D) 3, 2, 1

1.2 Measure the rate of a chemical reaction, using the procedure that they have developed.

KEY CONCEPTS

The **rate** of a chemical reaction is a ratio of the amount of change in a substance to the amount of time required for that change to take place. The substance could be a reactant or product. The amount of change can be measured in mass units (grams), partial pressure units (kilopascals), concentration units (mol/L or ppm) or number of particles (moles). The time can be measured in seconds, minutes, years or any other appropriate unit.

Note: The rate of a reaction is <u>not</u> the same as the time that it takes for a reaction to occur.

Example: The rate at which a candle burns could be measured in grams of parafin consumed per minute. However, stating that the candle burned for 2 hours before going out does not express the rate of the reaction.

The definition of the rate of a reaction can be used to solve numerical problems involving a single reaction rate as well as problems relating the rate of change of one substance to the rate of change of another.

Example 1:

Magnesium reacts in hydrochloric acid to produce hydrogen gas and magnesium chloride. A student places a piece of magnesium into a hydrochloric acid solution and measures the volume of hydrogen gas produced at the end of each minute for ten minutes. If he reported that 12 mL of hydrogen gas were present after 2.0 minutes and that 36 mL of hydrogen gas were present after 6.0 minutes, find the average rate of production of hydrogen gas between 2.0 and 6.0 minutes.

Solution:

The change in the volume of hydrogen gas during this interval = 36 mL - 12 mL
= 24 mL

The time required = 6.0 minutes − 2.0 minutes = 4.0 minutes

Therefore, the rate = change in amount ÷ time = 24 mL ÷ 4.0 minutes
= 6.0 mL/minute

Example 2:

Consider the following reaction: $H_{2}(g) + I_{2}(g) \rightarrow 2\,HI(g)$
A scientist places 5.0 moles of hydrogen gas (H_2) and 5.0 moles of iodine gas, I_2, into a 2.0 litre container and measures the amount of hydrogen iodide, HI, present at the end of each minute. His data is given in the table below.

Time (minutes)	0.0	1.0	2.0	3.0	4.0	5.0	6.0
Moles of HI present	0.0	4.0	6.0	7.4	8.4	9.0	9.2

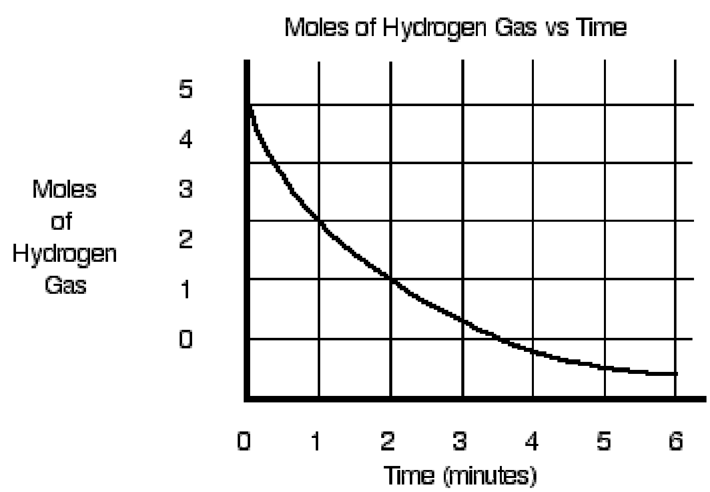
Sketch the corresponding graph of moles of hydrogen gas present vs time.

Solution:

Since we are working with moles and not concentration, the volume of the container is not relevant.
Knowing the amount of hydrogen iodide that had formed at the end of each time interval, we can find the amount of hydrogen gas that must have reacted. Remember that two moles of hydrogen iodide were formed when one mole of hydrogen gas reacted.
At 1.0 minute, 4.0 moles of hydrogen iodide were present. Therefore 2.0 moles of hydrogen gas must have reacted. Since 5.0 moles were present initially, there must have been 3.0 moles of hydrogen gas at 1.0 minute.
At 2.0 minutes, 6.0 moles of hydrogen iodide were present. Therefore 3.0 moles of hydrogen gas must have reacted. Since 5.0 moles were present
initially, there must have been 2.0 moles of hydrogen gas at 2.0 minutes.
Continuing in the same manner produces the following table.

Time (minutes)	0.0	1.0	2.0	3.0	4.0	5.0	6.0
Moles of hydrogen gas	5.0	3.0	2.0	1.3	0.8	0.5	0.4

This produces the following graph



SAMPLE QUESTIONS

1. *The conversion of sulfur dioxide to sulfur trioxide is represented by the following equation:* $2\ SO_{2(g)} + O_{2(g)} \rightarrow 2\ SO_{3(g)} + 206\ kJ$

   Which of the following could be used to define the rate of this reaction?

   A) The time required to produce one mole of sulfur trioxide
   B) The number of moles of sulfur trioxide produced per minute
   C) The number of moles of oxygen gas consumed
   D) The quantity of energy released per mole of product formed

2. *Consider the following reaction:* $N_{2(g)} + 3\ H_{2(g)} \rightarrow 2\ NH_{3(g)}$
   Under a certain set of conditions, the rate of formation of ammonia, (NH_3), was found to be 24 litres per minute. At what rate was hydrogen, (H_2), being consumed ?

   A) 12 litres per minute
   B) 16 litres per minute
   C) 24 litres per minute
   D) 36 litres per minute

3. *The electrolysis of water produces hydrogen gas according to the following equation.*

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

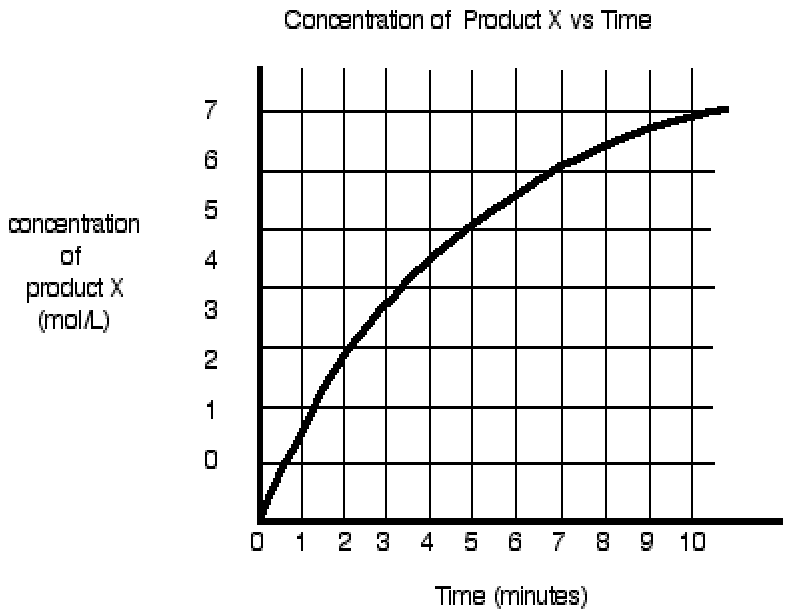
   A scientist must produce 24.0 g of oxygen using an apparatus that decomposes 45.0 mL of water per hour at room conditions. How much time is required to produce the required amount of oxygen gas?

   A) 18 minutes
   B) 36 minutes
   C) 72 minutes
   D) 96 minutes

4. *Experimental rate data was collected for the following hypothetical reaction:*

$$2\,XY \rightarrow 2\,X + Y_2$$

at constant temperature. The graph below expresses the concentration of product X as a function of time.



What was the average rate of formation of product Y_2 in the first 8 minutes? Express the answer in moles per litre per minute. ***(Show all necessary work.)***

2.1 Identify the factors that can influence the rate of a combustion reaction, after doing an experiment according to a suggested procedure.

KEY CONCEPTS

Combustion: the reaction of a chemical (fuel) with oxygen to produce oxides and heat

Combustion requires a fuel, a source of oxygen and a sufficiently high temperature known as the kindling point. Fires are extinguished by removing one of these components (or separating them).
Combustion rates are affected by the nature of the fuel used, the area of the fuel, the concentration of oxygen, the temperature and the presence of catalysts.
(The next section discusses factors in more detail.)

SAMPLE QUESTIONS

1. *Which of the following is not necessary for combustion to take place?*

   A) a fuel
   B) a source of oxygen
   C) a source of spark or flame
   D) a sufficiently high temperature

2. *Coal burns in air to produce carbon dioxide and water vapor. Under which of the following conditions would you expect the fastest reaction rate?*

   A) Chunks of coal are heated and allowed to burn naturally in air.
   B) Chunks of coal are allowed to burn while heated air is forced up through them.
   C) Coal is powdered and allowed to burn naturally in air.
   D) Coal is powdered and allowed to burn while heated air is forced up through it.

3. *Some fire extinguishers spray a foam over a burning substance to stop the fire. Why is the foam able to extinguish the fire?*

   A) The foam cools the surface to below its kindling point.
   B) The foam prevents oxygen from getting to the substance.
   C) The foam reacts with oxygen and totally depletes the oxygen supply.
   D) The foam contains an oxidizing agent that consumes the fuel without combustion.

2.2 Determine the effect of factors that influence the rate of a chemical reaction, after doing experiments according to a suggested procedure.

KEY CONCEPTS

The following factors are known to affect reaction rates.

1. The nature of the reacting substances: As a general rule, reactions involving ionic bond changes are usually more rapid than reactions involving covalent bond changes. Precipitation reactions are always rapid.

2. The concentration of the substances (in aqueous solutions or gases): Increasing the concentration of the reactants usually increases the reaction rate.

3. The surface area in contact: Increasing the surface area of the reacting substances usually increases the reaction rate.

4. The presence of a catalyst: A catalyst increases the rate of a chemical reaction.

5. The temperature: Increasing the temperature increases the rate of the reaction.

    Note: Changing the pressure on a gaseous system changes the concentration and therefore affects rate. Changing the pressure on systems that do not involve gases will not affect the rate.

A **Catalyst** is a substance that speeds up a chemical reaction without itself being consumed during the reaction.

An **Inhibitor** is a substance that can slow or stop a reaction when present in a reaction system.

SAMPLE QUESTIONS

1. *Consider the following system:*

$$Fe^{3+}{}_{(aq)} + SCN^{-}{}_{(aq)} \rightarrow FeSCN^{2+}{}_{(aq)}$$

Which of the following factors has no effect on the rate of reaction of this system?

A) A change of pressure
B) A change in the concentration of one of the reactants
C) The addition of a catalyst
D) A change of temperature

2. *Pieces of magnesium are reacted in test tubes containing hydrochloric acid at different concentrations and at different temperatures. Select the combination of acid concentration and temperature that would probably produce the greatest rate.*

   A) 1.0 mol/L at 12°C
   B) 3.0 mol/L at 12°C
   C) 1.0 mol/L at 36°C
   D) 3.0 mol/L at 36°C

3. *Based on the number and type of bonds to be broken and formed, which of the following chemical reactions would you expect to occur at the fastest rate at room temperature?*

   A) $CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)}$

   B) $Ag^{+}_{(aq)} + F^{-}_{(aq)} \rightarrow AgF_{(s)}$

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

   D) $Fe_{(s)} + S_{(s)} \rightarrow FeS_{(s)}$

2.3 Illustrate the effect of factors that influence the rate of a chemical reaction, using an analogical model.

KEY CONCEPTS

Collision theory is the idea that chemical reactions proceed when the reacting molecules collide with sufficient energy to result in the rearrangement of the atoms (an effective collision).

The **Activation energy** is the minimum energy with which particles must collide in order for the collision to be effective (result in the rearrangement of the bonds to form a new substance).

The **Activated complex** is the temporary, unstable arrangement of particles present at the highest potential energy point in a chemical reaction step.

The temperature of a sample of matter reflects the average amount of kinetic energy that the molecules possess. However, not every molecule in the sample possesses the same amount of kinetic energy. The distribution of Kinetic Energies graphs that follow show the relative number of molecules that possess different amounts of energy.

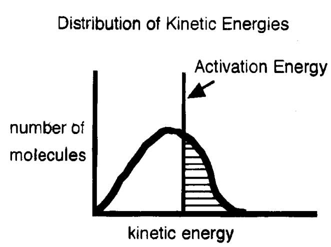


Diagram A

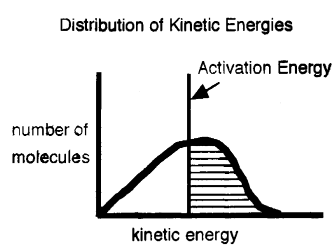


Diagram B
(increasing the temperature)

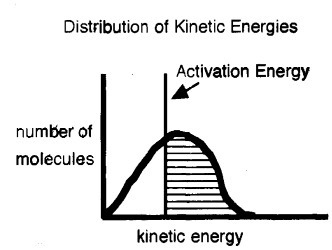
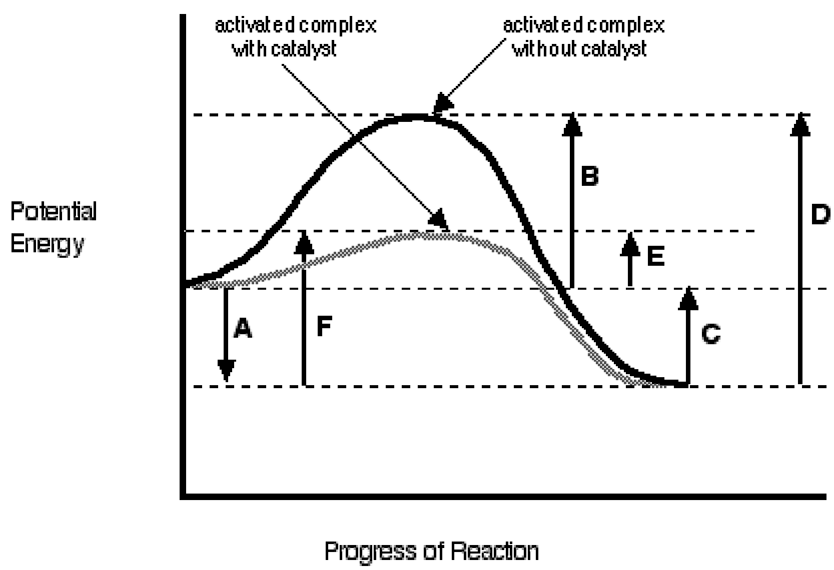


Diagram C
(Adding a catalyst)

The shaded area under the curve to the right of the activation energy line represents those molecules with sufficient kinetic energy to react. The greater the area to the right of the line, the faster the reaction will proceed.
An increase in temperature will change the shape of the curve so that the highest point is moved to the right, resulting in a greater shaded area. This indicates that more molecules have sufficient energy to react. The reaction will proceed at a greater rate. (see Diagram B)
Addition of a catalyst lowers the activation energy which shifts the vertical activation energy line to the left. The result is a greater area to the right of the line and therefore a greater reaction rate. (see Diagram C)

As molecules collide, their kinetic energy is changed to potential energy as the bonds between the atoms are stretched. If the collision is effective, the bonds break, new bonds form and the potential energy is changed back into kinetic energy as the resulting molecules separate. A graph of the potential energy of the system of reacting molecules vs the progress of the reaction is shown below. The black curve represents the reaction without a catalyst. The grey curve represents the possible position of the curve for the same reaction in the presence of a catalyst. Note that the presence of the catalyst lowers the activation energy of both the forward and reverse reactions but does not change the heat of reaction (ΔH).

Potential Energy Graph for the Reaction $X + Y \rightarrow XY$



The arrow A represents the change in enthalpy (ΔH) of the forward reaction.

The arrow B represents the activation energy (E_A) of the uncatalysed forward reaction.

The arrow C represents the change in enthalpy (ΔH) of the reverse reaction.

The arrow D represents the activation energy (E_A) of the uncatalysed reverse reaction.

The arrow E represents the activation energy (E_A) of the catalysed forward reaction.

The arrow F represents the activation energy (E_A) of the catalysed reverse reaction.

Each of the factors affecting reaction rate that were discussed in Objective 2.2 can be explained using collision theory.

The nature of the reacting substances determines the strength of the bonds in the reacting substances. The stronger the bonds, the higher the activation energy and the slower the reaction rate.

A higher concentration of reactants results in more collisions between the molecules, which results in a greater reaction rate.

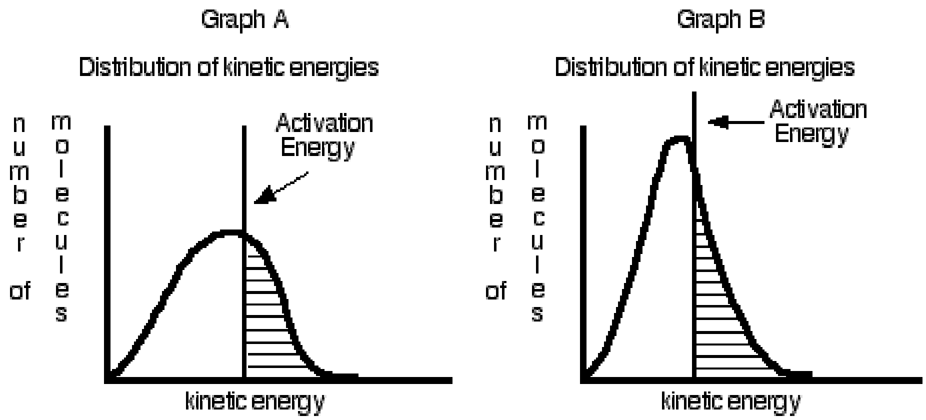
A greater surface area also results in more collisions between the molecules and a greater reaction rate.

Raising the temperature of the molecules causes the molecules to move faster, which results in more collisions. It also means that the molecules have more kinetic energy, which makes the collisions more likely to be effective. Therefore increasing the temperature increases the reaction rate.

A catalyst provides a reaction pathway with a lower activation energy, which makes the collisions more likely to be effective. Therefore, the presence of a catalyst increases the reaction rate.

SAMPLE QUESTIONS

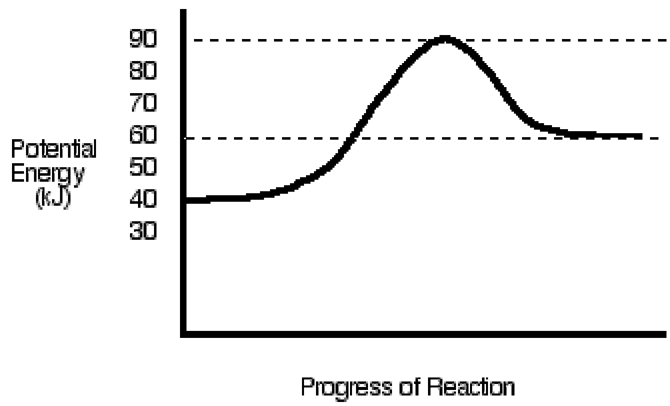
1. *Graph A represents an energy distribution graph for a chemical reaction. Graph B shows an energy distribution graph for the same reaction after one or more changes to the existing conditions were made.*



Which of the following changes could account for the differences in the two graphs?

A) The temperature was raised and a catalyst was added.
B) The temperature was lowered and a catalyst was added.
C) The temperature was lowered and a catalyst removed.
D) The temperature was lowered but no change in the catalyst was made.

2. *Consider the following graph of potential energy vs the progress of a reaction.*



From this graph, determine the activation energy, E_A for the forward reaction and the heat of reaction (ΔH) for the reverse reaction.

A) E_A = 50 kJ, ΔH = 20 kJ
B) E_A = 90 kJ, ΔH = -20 kJ
C) E_A = 50 kJ, ΔH = -20 kJ
D) E_A = 90 kJ, ΔH = 20 kJ

3. *A catalyst influences the rate of a chemical reaction by ______________*

A) increasing the potential energy of the reacting molecules
B) lowering the activation energy of the forward reaction while increasing the activation energy of the reverse reaction
C) increasing the temperature of the system
D) lowering the activation energy of both the forward and reverse reactions

4. *Which of the following changes increases the rate of a chemical reaction by increasing the number of collisions between the reacting molecules?*

1) The addition of a catalyst
2) An increase in temperature
3) An increase in the concentration of one of the reactants

A) 1 and 2 only
B) 1 and 3 only
C) 2 and 3 only
D) 1, 2 and 3

3.1 Illustrate, after consulting reference materials, the cause-and-effect relationships between society and the development and use of knowledge relating to the rate of chemical reactions, at different periods of history.

3.3 Illustrate the cause-and-effect relationships between the development and use of knowledge relating to the rate of chemical reactions and the natural and artificial environment, based on observations of phenomena and using reference materials.

KEY CONCEPTS

The study of reaction rates has led scientists to a better understanding of the way in which reaction rates affect natural processes and the development of technologies to control reaction rates. This knowledge has affected our society, the economy and the environment.

Examples of reaction rates playing a role in natural processes are:
Chlorophyll acting as a catalyst in photosynthesis

Enzymes (protein molecules that act as catalysts in organic reactions) in the body that allow glucose to undergo oxidation rapidly at relatively low temperatures

The slowing of body processes at lower temperatures which reduces the need for food and oxygen during the hibernation of animals

Examples of technology that uses the knowledge of reaction rates are:
The use of refrigeration to slow the rate at which food spoils
The use of preservatives (chemical inhibitors) to prevent food spoilage
The use of low-temperature surgery to slow body processes during the operation to reduce blood loss and the need for oxygen
The use of catalytic converters in automobiles to reduce pollution by increasing the rate of oxidation or reduction of various exhaust gases
The use of enzymes in laundry detergents to remove clothing stains

SAMPLE QUESTION

1. *Which of the following technologies does not owe its effectiveness primarily to the alteration of rates of reaction?*

   A) catalytic converters on automobile exhaust systems to reduce pollution
   B) refrigeration to prevent food spoilage
   C) genetic alteration of plants to produce larger tomatoes
   D) chemical fire extinguishers to put out fires

MODULE 5
- EQUILIBRIUM IN CHEMICAL REACTIONS -

1.1	**Based on observations of the macroscopic behavior of matter, distinguish between a chemical reaction occurring in a closed environment and one occurring in an open environment.**
1.2	**Determine the effect of various factors on the equilibrium of chemical system, based on experimental evidence.**
2.1	**Represent the dynamic microscopic aspect of a balanced system so as to understand how it functions.**
2.2	**Establish a mathematical equation using the concentrations of products and reactants in a balanced chemical reaction, at constant temperature.**

KEY CONCEPTS

Equilibrium

The term "equilibrium" is a term that implies a balancing of some kind. The equilibrium state is found around us in everyday life. Equilibrium may be physical or chemical. It may also be static or dynamic.

A person sitting on a chair with the four legs on the floor is in a state of equilibrium. The moon rotating around the earth is in a state of equilibrium. As a person walks, he/she goes through various states of equilibrium. The laws of Physics show many examples of equilibrium. Both static (stationary, as in sitting on a chair) and dynamic (as in the moon's rotation and a person walking) are possible.

Chemical reactions are generally reversible. If there is to be a "balancing" of the system, then the rates of the forward and reverse chemical changes MUST be equal.

For the general reaction:

$$aA + bB \rightarrow cC + dD$$

where the substances A, B, C and D can exist in variable concentrations, the **Law of Mass Action** states that the rate of the forward reaction is related to the concentrations of the reacting substances by the following formula:

$$Rate_f = k_f\,[A]^a[B]^b$$

the rate of the reverse chemical reaction would be:

$$Rate_r = k_r[C]^c[D]^d$$

For equilibrium to occur, $Rate_f = Rate_r$.

$$k_f[A]^a[B]^b = k_f[C]^c[D]^d$$

$$\frac{k_f}{k_r} = \frac{[C]^c[D]^d}{[A]^a[B]^b} \qquad\qquad \frac{k_f}{k_r} = \boldsymbol{K}_c$$

$$\text{Equilibrium constant} = \boldsymbol{K}_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

This relationship is known as the **Equilibrium Law** or the **Equilibrium Constant Expression**.

The larger the value of $\boldsymbol{K_c}$, the more the reaction favors the products at equilibrium. When using both the Law of Mass Action and the Equilibrium Law, we use only those substances whose concentrations can vary. Because the concentration of a pure liquid or solid does not vary, liquids and solids do not appear in equilibrium constant expressions. In other words, include only substances in the gaseous or aqueous states when using the Equilibrium Law.

Sometimes a reaction involves only gases. In this case, the concentration of each gas is proportional to its partial pressure. Consider the following reaction:

$$aA_{(g)} + bB_{(g)} \rightarrow cC_{(g)} + dD_{(g)}$$

We can express the equilibrium constant as:

$$K_p = \frac{(P_C)^c(P_D)^d}{(P_A)^a(P_B)^b}$$

When evaluating an equilibrium constant, **K_c**, concentrations are expressed in mol/L. When evaluating an equilibrium constant, **K_p**, pressures are expressed in kPa. When the value of an equilibrium constant is stated, the units are not included, but the temperature is stated because a change in temperature will change the value of the equilibrium constant.

Some examples of chemical reactions and their equilibrium law expressions are given below:

1: $N_{2(g)} + 3\,H_{2(g)} \leftrightarrow 2\,NH_{3(g)}$

$$K_c = \frac{[NH_3]^2}{[N_2]\,[H_2]^3} \quad \text{or} \quad K_p = \frac{(P_{NH_3})^2}{(P_{N_2})\,(P_{H_2})^3}$$

2: $MgCl_{2(s)} + H_2O_{(l)} \leftrightarrow Mg^{2+}\,(aq) + 2\,Cl^-\,(aq)$

$$K_c = [Mg^{2+}][Cl^-]^2$$

3: $Zn_{(s)} + 2\,Ag^+\,(aq) \leftrightarrow 2\,Ag_{(s)} + Zn^{2+}\,(aq)$

$$K_c = \frac{[Zn^{2+}]}{[Ag^+]^2}$$

A chemical system is at **EQUILIBRIUM** if it meets the four following criteria:

1. The system is a closed system. (The reaction is in a sealed container or can be considered closed for a specific period of time.)

2. The reaction is at constant temperature.

3. There is no macroscopic activity. (Nothing seems to be happening; the properties of the system are constant.)**

4. There is molecular activity. (Something is happening on the molecular level.)

**** These constant properties include:**

- **color**
- **excess solute (undissolved solid)**
- **constant concentration of each substance present in the reaction**
- **pressure**

<u>Equilibrium vs Steady State</u>

A system may seem to be at equilibrium but does not meet all the conditions above. This situation is called a "steady state".

The following examples can be used to distinguish between equilibrium and a steady state:

A stopppered flask contains some water but is not full. The water level is constant. Since the rate of evaporation is equal to the rate of condensation, this is san equilibrium system.

A dog is kept in a cage and given food and water. His mass stays constant. Since the system is not closed and the process is not reversible, this is a steady state system.

A candle is burning. The temperature and size of the flame are constant. Since this system is not closed and the products of the combustion are not being recycled, there is no reverse process. This is a steady state system.

Nitrogen gas and hydrogen gas are reacting to produce ammonia gas in a sealed container. Eventually, the concentrations of the nitrogen, hydrogen and ammonia remain constant. Because all three gases are present and maintain constant concentrations, the rate of formation of ammonia must be equal to the rate of decomposition of ammonia. This is an equilibrium system.

A student prepares a solution of sodium chloride with excess solid sitting on the bottom of the flask. He covers the flask. The next day, he notices that the amount of solid on the bottom has remained constant. Since the amount of solid remained constant, the rate of dissolving of the solid must be equal to the rate of precipitation. This is an equilibrium system.

A car runs into a fire hydrant. The water shoots into the air from the water line, maintaining a constant height and rate of flow. Since the water that flows from the line leaves the system and is replaced by different water, this system is not closed. This is a steady state system.

Effect of a Catalyst

Chemical equilibrium will be attained in all reversible chemical reactions at some point in time. The time required to reach equilibrium will depend on the specific reaction. Some reactions will reach equilibrium almost instantaneously; these are generally ionic reactions. Other reactions may take long periods of time before equilibrium is attained. The addition of a catalyst to a system already at equilibrium will have NO effect. The role of a catalyst (as explained previously) is to alter the RATE of a reaction. If a catalyst is added to a reversible chemical reaction at the START of the reaction, it will alter the time to reach equilibrium.

Mathematical Application of the Law of Mass Action

When an equilibrium constant is to be calculated, equilibrium concentrations of the reactants and products must be used. If these are not given, other suitable information must be used to find the equilibrium concentrations. Some examples are given below:

Example 1:

At a given temperature , the concentrations of the following gases are given at equilibrium:

$[CO] = 0.2$ mol/L, $[H_2O] = 0.5$ mol/L, $[H_2] = 0.32$ mol/L, and
$[CO_2] = 0.42$ mol/L.

Calculate the equilibrium constant given the equation

$$CO(g) + H_2O(g) \leftrightarrow H_2(g) + CO_2(g)$$

Solution

$$K_c = \frac{[H_2]\,[CO_2]}{[CO]\,[H_2O]} = \frac{0.32 \times 0.42}{0.2 \times 0.5} = 1.34$$

Note: Units are not generally given with equilibrium constants primarily because there are too many possibilities.

Example 2:

What is the equilibrium concentration of SO_3 in the following reaction if the equilibrium concentrations of SO_2 and O_2 are each 0.05 mol/L and $\boldsymbol{K}_c = 85.0$?

Solution

$$2\ SO_{2(g)} + O_{2(g)} \leftrightarrow 2\ SO_{3(g)}$$

$$\boldsymbol{K}_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

$$85.0 = \frac{[SO_3]^2}{(0.05)^2 \times 0.05}$$

$$[SO_3]^2 = 85.0 \times (0.05)^2 \times 0.05$$

$$= 0.0106$$

$$[SO_3] = 0.103 \text{ mol/L}$$

Example 3

Given the reaction:

$$H_2(g) + I_2(g) \leftrightarrow 2\,HI(g)$$

Initially, 4 moles of hydrogen gas and 3 moles of iodine gas are put in a 5 L container. At equilibrium, the concentration of hydrogen iodide is 0.3 mol/L. Find the equilibrium constant. (eqm as used below means equilibrium)

Solution

initial $[H_2]$ = 4 mol/ 5 L = 0.8 mol/L

initial $[I_2]$ = 3 mol/ 5 L = 0.6 mol/L

eqm [HI] = 0.3 mol/L

eqm $[H_2]$ = initial $[H_2]$ – amount reacted
= 0.8 mol/L – 0.15 mol/L
= 0.65 mol/L

eqm $[I_2]$ = initial $[I_2]$ – amount reacted
= 0.6 mol/L – 0.15 mol/L
= 0.45 mol/L

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(0.3)^2}{(0.65)(0.45)} = 0.31$$

NOTE: 2 mol of HI comes from 1 mol of H_2 and 1 mol of I_2

Another way of solving the first part of *example 3* is the following:

	$H_2(g)$	+	$I_2(g)$	$\leftrightarrow$	2 HI(g)
I	0.8		0.6		0
C	-0.15		-0.15		+0.3
E	0.65		0.45		0.3

WHERE:
I is the initial condition
C is the change (shift) (reacteD)
E is the equilibrium condition

The remaining part of the problem is solved as above.

SAMPLE QUESTIONS

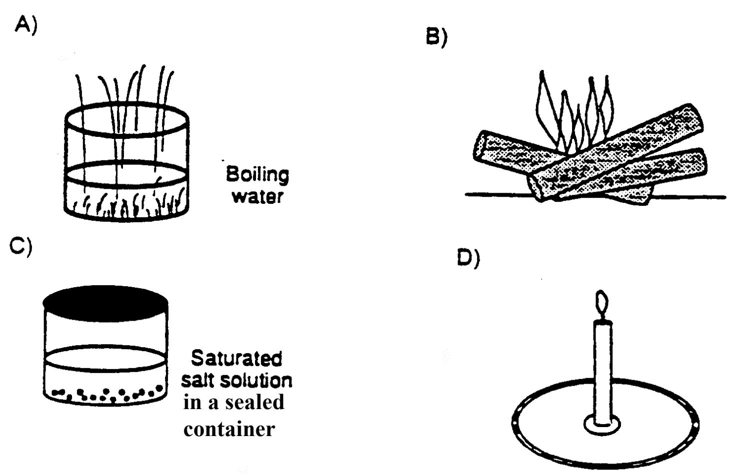
1. *Which of the following is not a property of a system at equilibrium?*

   A) The system is closed.
   B) The reactants are completely transformed into products.
   C) The temperature is constant.
   D) The reaction is reversible.

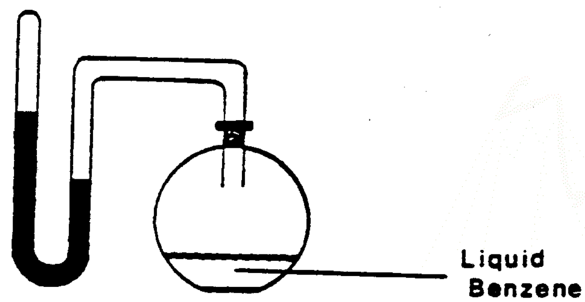
2. *Which of the following statements defines the dynamic nature of a chemical equilibrium?*

   A) The reactants transform completely into products.
   B) The macroscopic properties remain constant.
   C) The masses of the reactants and the products are equal.
   D) The rates of the forward and reverse reactions are equal.

3. *Which of the following systems illustrated below is in a state of equilibrium?*



4. *The diagram below represents a flask containing liquid benzene attached to a manometer.*



After a certain period of time, the manometer indicates that the vapor pressure of the benzene is CONSTANT. Identify the statement which explains this fact.

A) After a certain time, the liquid benzene stops evaporating.
B) After a certain time, all of the liquid benzene has evaporated.
C) After a certain time, the rate of evaporation of liquid benzene equals the rate of condensation of gaseous benzene.
D) After a certain time, the vapor pressure of benzene equals 100 kPa.

5. *A change in which of the following factors can affect a system at equilibrium?*

1) Temperature
2) Pressure
3) Catalyst
4) Concentration
5) Surface area of contact

A) 1, 2 and 3
B) 1, 2 and 4
C) 1, 4 and 5
D) 2, 3 and 5

6. *Which statement concerning the following equilibrium systems is FALSE?*

1) $N_{2(g)} + 3\ H_{2(g)} \leftrightarrow 2\ NH_{3(g)}$ $K_c = 2.66 \times 10^{-3}$

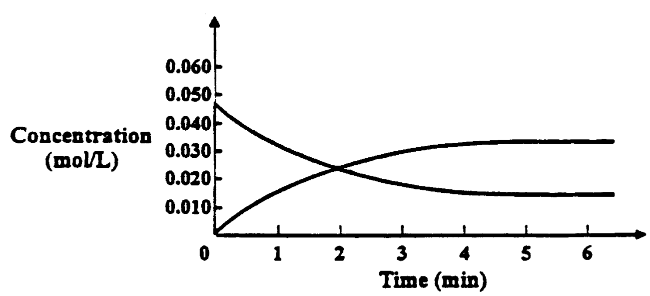
2) $2\ H_{2(g)} + S_{2(g)} \leftrightarrow 2\ H_2S_{(g)}$ $K_c = 9.38 \times 10^{-5}$

3) $2\ H_2O_{(g)} + 2\ S_{(s)} \leftrightarrow 2\ H_2S_{(g)} + O_{2(g)}$ $K_c = 5.31 \times 10^{-10}$

4) $H_{2(g)} + I_{2(g)} \leftrightarrow 2\ HI_{(g)}$ $K_c = 54.4$

A) The formation of $NH_{3(g)}$ is favored in system 1.
B) The formation of HI(g) is favored in system 4.
C) The formation of H_2S is not favored in system 2.
D) The formation of water is favored in system 3.

7. *The following is a graph of the concentration of products and reactants as a function of time as a reaction proceeds.*



Which of the following times corresponds to the FIRST point at which equilibrium has been reached?

A) 2 min
B) 3 min
C) 5 min
D) 6 min

8. *Which of the following changes will increase the numerical value of the equilibrium constant?*

A) decreasing the temperature of an equilibrium system in which the forward reaction is exothermic
B) increasing the concentrations of the reactants in an equilibrium system
C) increasing the pressure on an equilibrium system involving only gases
D) adding a catalyst to an equilibrium system

9. *What is the mathematical expression for the equilibrium constant for the reaction represented below?*

$CH_{4(g)} + 2\ O_{2(g)} \leftrightarrow CO_{2(g)} + 2\ H_2O_{(g)}$

A) $K_c = \dfrac{[CO_{2(g)}][H_2O_{(g)}]^2}{[CH_{4(g)}][O_{2(g)}]^2}$

B) $K_c = \dfrac{[CO_{2(g)}]}{[CH4(g)[O2(g)]^2}$

C) $K_c = \dfrac{[CH_{4(g)}][O_{2(g)}]^2}{[CO_{2(g)}][H_2O_{(g)}]^2}$

D) $K_c = \dfrac{[CH_{4(g)}][O_{2(g)}]^2}{[CO_{2(g)}]}$

10. *What is the correct expression for the equilibrium constant for the reaction?*

$CN^-\ (aq) + H_2O_{(l)} \leftrightarrow HCN_{(aq)} + OH^-\ (aq)$

A) $K_c = \dfrac{[HCN_{(aq)}][OH^-\ (aq)]}{[CN^-\ (aq)][H_2O_{(l)}]}$

B) $K_c = \dfrac{[OH^-\ (aq)]}{[CN^-\ (aq)][H_2O_{(l)}]}$

C) $K_c = \dfrac{[HCN_{(aq)}][OH^-\ (aq)]}{[CN^-\ (aq)]}$

D) $K_c = \dfrac{[CN^-\ (aq)]}{[HCN_{(aq)}][OH^-\ (aq)]}$

11. *In studying the equilibrium of a chemical system, a student observed that the behavior of the chromate ion, CrO_4^{2-} (aq), depends on the acidity of the system. When the acidity of the system changes, the color of the solution can vary from yellow to orange. This reaction is illustrated by the following equation:*

   $2\ CrO_4^{2-}(aq) + 2\ H^+(aq) \leftrightarrow Cr_2O_7^{2-}(aq) + H_2O(l)$

   Which expression should be used to find the equilibrium constant, K_c, of this system?

   A) $K_c = \dfrac{[CrO_4^{2-}]^2[H^+]^2}{[Cr_2O_7^{-2}]}$

   B) $K_c = \dfrac{[Cr_2O_7^{2-}][H_2O]}{[CrO_4^{2-}]^2\ [H^+]^2}$

   C) $K_c = \dfrac{[CrO_4^{2-}]^2\ [H^+]^2}{[Cr_2O_7^{2-}][H_2O]}$

   D) $K_c = \dfrac{[Cr_2O_7^{2-}]}{[CrO_4^{2-}]^2\ [H^+]^2}$

12. *What is the mathematical expression for K_c of the equilibrium system represented by the following equation?*

   $Cu_{(s)} + 2\ H_2SO_{4(aq)} \leftrightarrow CuSO_{4(aq)} + SO_{2(aq)} + 2\ H_2O_{(l)}$

   A) $K_c = \dfrac{[CuSO_4][SO_2][H_2O]^2}{[Cu][H_2SO_4]^2}$

   B) $K_c = \dfrac{[CuSO_4][SO_2][H_2O]^2}{[H_2SO_4]^2}$

   C) $K_c = \dfrac{[CuSO_4][SO_2]}{[H_2SO_4]^2}$

   D) $K_c = [SO_2]$

13. *For the following system, the equilibrium constant has a value of 1.30 at 25^oC. If the temperature is increased to 250^oC, which of the following statements concerning the equilibrium constant is TRUE?*

$CO_{2(g)} + H_{2(g)} + Energy \leftrightarrow CO_{(g)} + H_2O_{(g)}$

A) Its value will be equal to 1.30.
B) Its value will be greater than 1.30.
C) Its value will be less than 1.30.
D) Its value will be 13.0.

14. *What is the equilibrium constant for the following system if at equilibrium there are 3.0 mol/L of $NO_{2(g)}$ and 4.0 mol/L of $N_2O_{4(g)}$?*

$2\ NO_{2(g)} \leftrightarrow N_2O_{4(g)} + Energy$

A) 0.44
B) 1.30
C) 0.75
D) 2.30

15. *In a closed system, the initial $[N_2O_{4(g)}]$ is 2.0 mol/L. At equilibrium the $[NO_{2(g)}]$ is 0.80 mol/L.*

$N_2O_{4(g)} \leftrightarrow 2\ NO_{2(g)}$

Calculate the K_c value of this system.

A) 0.16
B) 0.40
C) 0.50
D) 1.60

16. *Given the following system:*

$2\ NO_{2(g)} \leftrightarrow 2\ NO_{(g)} + O_{2(g)}$

5.0 moles of $NO_{2(g)}$ are placed in a 1 litre container. At equilibrium, there are 1.5 moles of $O_{2(g)}$ present. Calculate the value of K_c.

A) 0.54
B) 2.20
C) 3.38
D) 6.70

17. A student places 4.0 mol of A and 4.0 mol of B in a 1 litre flask. When equilibrium is reached, there are 2.0 mol of X. Calculate the K_c for this system.

A(g) + 2 B(g) $\leftrightarrow$ 2 X(g) + 4 Y(g)

A) 1.30
B) 5.30
C) 16
D) 85

18. A student adds 3.0 moles of $N_{2(g)}$ and 6.0 moles of $O_{2(g)}$ to a 5.0 L container. At equilibrium 1.0 mole of $NO_{2(g)}$ is present. Calculate the equilibrium constant for this system.

$N_{2(g)} + 2\ O_{2(g)} \leftrightarrow 2\ NO_{2(g)}$

A) 1.0×10^{-3}
B) 9.2×10^{-3}
C) 8.0×10^{-2}
D) 4.0×10^{-1}

19. Given the following equilibrium constant, choose the correct chemical reaction for this formula.

$$K_c = \frac{[CO_{2(g)}]^3\ [H_2O_{(g)}]^4}{[C_3H_{8(g)}][O_{2(g)}]^5}$$

The reaction is:
A) $3\ CO_{2(g)} + 4\ H_2O_{(g)} \leftrightarrow C_3H_{8(g)} + 5\ O_{2(g)}$
B) $C_3H_{8(g)} + O_{2(g)} \leftrightarrow 3\ CO_{2(g)} + 4\ H_2O_{(g)}$
C) $3\ CO_{2(g)} + H_2O_{(g)} \leftrightarrow C_3H_{8(g)} + 5\ O_{2(g)}$
D) $C_3H_{8(g)} + 5\ O_{2(g)} \leftrightarrow 3\ CO_{2(g)} + 4\ H_2O_{(g)}$

20. *A student heated a mixture of sulfur dioxide, SO_2, and oxygen gas, O_2, in a 10.0L container.*

The reaction is represented by the following equation:

$$2\ SO_{2(g)} + O_{2(g)} \leftrightarrow 2\ SO_{3(g)} + \text{Energy}$$

At a temperature of $70^{o}C$, the reaction reached equilibrium and the composition of the mixture was as follows:

$SO_{2(g)}$: 3.0 mol
$O_{2(g)}$: 0.5 mol
$SO_{3(g)}$: 1.5 mol

What is the equilibrium constant, K_c, for this reaction?

1.3 Predict the effect of various factors on the concentration of substances and on the direction of the shift in the equilibrium of a chemical system, based on Le Châtelier's Principle.

KEY CONCEPTS

Le Châtelier's Principle states that a system at equilibrium will stay at equilibrium UNLESS it is acted upon externally. If an imposed change acts on an equilibrium situation, the equilibrium will re-establish itself with new equilibrium conditions. These new conditions will be counteracting the external stress applied.

The different changes that can be made to a set of equilibrium conditions are:

i-increasing or decreasing the concentration of one substance
ii-increasing or decreasing the pressure for a gas reaction
iii-increasing or decreasing the temperature

Let us examine each of these changes and find the effect on the equilibrium for a general reaction:

$$aA + bB \leftrightarrow cC + dD$$

At equilibrium, there is a certain amount of A, B, C and D present. These amounts will remain unchanged UNLESS some external stress is applied.

i- increasing or decreasing the concentration of one substance:

Example 1: the addition of some B

The addition of more B means that there are more B molecules available to react. The initial effect of the addition of some B will be to drastically favor the forward process (reaction shifts to the right). After a period of time, the reverse process will also reestablish itself. When the new equilibrium is attained, there will be less A present, more B present, more C present, and more D present than at the original equilibrium.

Example 2: the removal of some A

The removal of some A will cause the reaction to want to replace the removed A molecules. The initial effect of the removal of some A will be to drastically favor the reverse process (reaction shifts to the left). After a period of time, the forward process will also reestablish itself. When the new equilibrium is attained, there will be less A present, more B present, less C present, and less D present than at the original equilibrium.

ii- increasing or decreasing the pressure

Since pressure for a gas depends on the number of gas particles present, an increase or decrease in pressure will have an effect on an equilibrium involving gases IF there are a different number of gas particles on each side of the balanced equation. For the general reaction, one must look at "a+b" and "c+d". If these are the same value, a change in pressure will have NO effect – there are the same number of gas particles on both sides of the equation.

An increase in pressure will favor the side with the fewer gas particles. A decrease in pressure will favor the side with the larger number of gas particles.

Example 1: What is the effect of increasing the pressure on the reaction:

$N_{2(g)} + 3\ H_{2(g)} \leftrightarrow 2\ NH_{3(g)}$?

Since there are 4 moles of reactant gases and 2 moles of product gas, an increase will cause the amount of NH_3 to increase and the amounts of H_2 and N_2 gases to decrease.

Example 2: What is the effect of decreasing the pressure on the reaction:

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

Since there are 2 moles of reactant gas and 2 moles of product gas, there will be no change in the amount of each substance present.

iii- increasing or decreasing the temperature

Previously you have learned that the term exothermic means that heat (energy) is released and it can be written into the equation as a product.

$A + B \leftrightarrow C + D + \text{Energy}$

For an exothermic reaction, an increase in temperature will favor the reverse process. The reasoning is the same as if energy/temperature was another substance in the equation. This makes the situation similar to section i -

increasing or decreasing the concentration of one substance. That is, the concentrations of the reactants will increase and the concentrations of the products will decrease. A decrease in temperature will have the opposite effect.

An endothermic reaction is one where heat (energy) is absorbed and it can be written into the equation as a reactant.

$$A + B + \text{Energy} \leftrightarrow C + D$$

For an endothermic reaction, an increase in temperature will favor the forward process. That is, the concentrations of the products will increase and the concentrations of the reactants will decrease. An decrease in temperature will have the opposite effect.

Example 1: ***What effect will there be for an increase in temperature applied to the following system:***

$$N_2O_4 + \text{energy} \leftrightarrow 2\ N_2O\ ?$$

An increase in energy means that there is more energy available to break the N_2O_4 bonds. This will result in less N_2O_4 and more NO_2 gas present at the new equilibrium.

Example 2: ***What effect will there be for an increase in temperature applied to the following system:***

$$N_2 + 3\ H_2 \leftrightarrow 2\ NH_3 + \text{energy}\ ?$$

An increase in energy means there is more energy available to break the NH_3 bonds. This will result in less NH_3 and more N_2 and H_2 at the new equilibrium.

Note*:* It is possible to predict the effect of more than one change applied to an equilibrium situation provided that these changes do not contradict each other.

Example: $N_{2(g)} + 3H_{2(g)} \leftrightarrow 2\ NH_{3(g)} + \text{energy}$

What conditions would produce a maximum yield of $NH_{3(g)}$?

To get the maximum amount of NH_3 possible, the ideal conditions would be at low temperature, high pressure, and high concentrations of hydrogen and nitrogen gases.

SAMPLE QUESTIONS

1. *Consider the following equilibrium equation:*

   $C_2H_{4(g)} + 3\ O_{2(g)} \leftrightarrow 2\ CO_{2(g)} + 2\ H_2O_{(g)} + Energy$

   Which of the following changes will favor the production of $CO_{2(g)}$*?*

   1. *lowering the temperature*
   2. *raising the pressure*
   3. *increasing the concentration of oxygen*
   4. *adding a catalyst*

   A) 1 and 2
   B) 1 and 3
   C) 2 and 4
   D) 3 and 4

2. *The following orange-yellow solution is a system at equilibrium:*

   $Cr_2O_7^{2-}{}_{(aq)} + 2\ OH^-{}_{(aq)} \leftrightarrow 2\ CrO_4^{2-}{}_{(aq)} + H_2O_{(l)}$
   orange colorless yellow colorless

   An acidic solution containing $H^{+1}_{(aq)}$ *ions is added to this system.*

   What will happen to this orange-yellow solution after equilibrium is re-established?

   A) It will become more orange.
   B) It will become more yellow.
   C) It will become colorless.
   D) It will show no change in color.

3. *The following table shows four systems at equilibrium and the change made to each system.*

SYSTEM AT EQUILIBRIUM	CHANGE MADE
1. $A_{(g)} + 2\ B_{(g)} \leftrightarrow C_{(g)}$	pressure was increased
2. $3\ D_{(g)} + 2\ E_{(g)} \leftrightarrow 2\ F_{(g)}$	a catalyst was added
3. $2\ E_{(g)} + G_{(g)} \leftrightarrow E_2G_{(g)}$	some $G_{(g)}$ was removed
4. $A_{(g)} + C_{(g)} \leftrightarrow 3\ D_{(g)}$	the volume was increased

In which system does the change favor the reverse reaction ONLY?

A) 1
B) 2
C) 3
D) 4

4. *By applying Le Châtelier's Principle to the following, identify the modifications, which favor the formation of products.*

$N_{2(g)} + 3\ H_{2(g)} \leftrightarrow 2\ NH_{3(g)}$ *+ Energy*

1) increasing the pressure
2) increasing the temperature
3) increasing the volume
4) adding a catalyst
5) increasing the concentration of $N_{2(g)}$

A) 1 and 5
B) 1, 2 and 4
C) 2, 3 and 4
D) 3 and 5

5. *Given the equilibrium system:*

$2\ CO_{(g)} + O_{2(g)} \leftrightarrow 2\ CO_{2(g)}$ *+ Energy*

What is the effect on the concentrations of $CO_{(g)}$ *and* $O_{2(g)}$ *if the concentration of* $CO_{2(g)}$ *is increased?*

A) $[CO_{(g)}]$ ***increases and*** $[O_{2(g)}]$ ***decreases***
B) $[CO_{(g)}]$ ***increases and*** $[O_{2(g)}]$ ***increases***
C) $[CO_{(g)}]$ ***decreases and*** $[O_{2(g)}]$ ***increases***
D) $[CO_{(g)}]$ ***decreases and*** $[O_{2(g)}]$ ***decreases***

6. *In the equilibrium system:*

   $4\ HCl_{(g)} + O_{2(g)} \leftrightarrow 2\ Cl_{2(g)} + H_2O_{(l)} + 113\ kJ$

   Which of the following operations can be used to increase the production of $Cl_{2(g)}$?

   A) increase the pressure and decrease the temperature
   B) increase the pressure and temperature
   C) decrease the pressure and increase the temperature
   D) decrease the pressure and temperature

7. *Given the following system at equilibrium:*

   $CO_{(g)} + 2\ H_{2(g)} \leftrightarrow CH_3OH_{(g)}$ + Energy

   Which of the following changes will increase the concentration of $CH_3OH_{(g)}$?

   A) increasing the concentration of $CO_{(g)}$
   B) decreasing the concentration of $H_{2(g)}$
   C) raising the temperature
   D) decreasing the pressure

8. *What will be the effect of increasing the pressure on the following equilibrium system?*

   $PCl_{5(g)} + 92.2\ kJ \leftrightarrow PCl_{3(g)} + Cl_{2(g)}$

   A) The decomposition of PCl_5 will be favored.
   B) The reactant will be favored.
   C) There will be no change.
   D) The temperature of the system will drop.

9. *What will be the effect of adding $Cu(NO_3)_{2(s)}$ to the equilibrium system represented by the following equation?*

   $Cu^{2+}{}_{(aq)} + 4\ NH_{3(g)} \leftrightarrow Cu(NH_3)_4{}^{2+}{}_{(aq)}$

   A) an increase in $[Cu^{2+}]$ and $[Cu(NH_3)_4{}^{2+}]$, and a decrease in $[NH_3]$
   B) an increase in $[Cu^{2+}]$, $[Cu(NH_3)_4{}^{2+}]$, and $[NH_3]$
   C) a decrease in $[Cu^{2+}]$ and $[NH_3]$ and an increase in $[Cu(NH_3)_4{}^{2+}]$
   D) no effect on $[Cu^{2+}]$, $[NH_3]$, and $[Cu(NH_3)_4{}^{2+}]$

10. *In which of the equilibrium systems represented by the following equations will an increase in pressure favor the forward reaction?*

    A) $CuBr_{2(s)} \leftrightarrow Cu_{(s)} + Br_{2(g)}$
    B) $H_{2(g)} + Cl_{2(g)} \leftrightarrow 2\ HCl_{(g)} + \text{Energy}$
    C) $CaCO_{3(s)} + \text{Energy} \leftrightarrow CaO_{(s)} + CO_{2(g)}$
    D) $2\ H_2O_{(l)} + O_{2(g)} \leftrightarrow 2\ H_2O_{2(l)}$

11. *Study the following system at equilibrium:*

    $N_{2(g)} + 2\ O_{2(g)} + 34\ kJ \leftrightarrow 2\ NO_{2(g)}$

    Which of the following factors listed below would result in the displacement of the equilibrium toward the right?

    a) an increase in pressure
    b) introducing a catalyst to the system
    c) an increase in the concentration of O_2
    d) cooling the system
    e) an increase in the concentration of NO_2
    f) an increase in the temperature

12. *Equilibrium is achieved in a closed system where metallic lead can react with hydrochloric acid. This system is represented by the following net ionic equation:*

    $Pb_{(s)} + 2\ H^{+}{}_{(aq)} \leftrightarrow Pb^{2+}{}_{(aq)} + H_{2(g)}$

    A sodium hydroxide pellet, NaOH(s), is added to this system. What happens to the concentration of each substance in the system?

13. *Thiocyanate anions,* $SCN^{-}{}_{(aq)}$*, react with Fe(III) cations,* $Fe^{3+}{}_{(aq)}$*, to form soluble red cations,* $Fe(SCN)^{2+}{}_{(aq)}$*, according to the equation:*

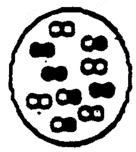
    1) $Fe^{3+}{}_{(aq)} + SCN^{-}{}_{(aq)} \leftrightarrow Fe(SCN)^{2+}{}_{(aq)}$
    light yellow colorless deep red

    If some silver nitrate, $AgNO_{3(aq)}$*, is added to system 1 above, it will cause the formation of a precipitate of* $AgSCN_{(s)}$ *as shown in the following equation:*

    2) $Ag^{+}{}_{(aq)} + SCN^{-}{}_{(aq)} \leftrightarrow AgSCN_{(s)}$

    Using Le Châtelier's Principle, discuss what will happen to the concentration of each substance in system 1 after the silver ions are added. Also discuss how this will affect the original colors present in system 1.

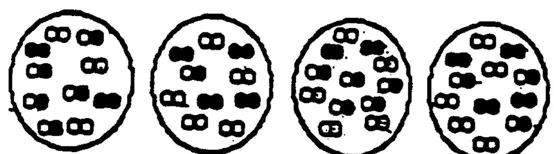
14. *Given an original equilibrium mixture below:*



A_2 + B_2 + Energy $\leftrightarrow$ 2 AB

●● + ○○ + *Energy* $\leftrightarrow$ 2 ●○ ●○

Which of the following diagrams best represents the new equilibrium mixture formed when the temperature of the system above is increased? Explain your reasoning.



A. B. C. D.

15. *The following equation represents the formation of hydrogen iodide, $HI_{(g)}$, from its elements:*

$H_{2(g)} + I_{2(g)} \leftrightarrow 2\ HI_{(g)} + 11\ kJ$

How will a temperature increase affect the value of the equilibrium constant for this system?

2.3 To calculate the value of the ionization constant of water at 25°C, based on the knowledge of equilibrium.

KEY CONCEPTS

Water is considered to be a non-electrolyte. However a few water molecules do react to form ions by the following equation:

$$H_20_{(L)} \rightarrow \underset{\text{hydrogen ions}}{H^+(aq)} + \underset{\text{hydroxide ions}}{OH^-(aq)}$$

since the pH of pure water at 25°C is 7, then the

Hydrogen ion concentration is $[H^+_{(aq)}] = 1 \times 10^{-7}$ mol/L

and

Hydroxide ion concentration is $[OH^-_{(aq)}] = 1 \times 10^{-7}$ mol/L

The equilibrium constant for water, at 25°C, is given by the following mathematical expression:

$$K_W = [H^+_{(aq)}]\,[OH^-_{(aq)}]$$

Therefore $$K_W = (1 \times 10^{-7} \text{ mol/L})(1 \times 10^{-7} \text{ mol/L})$$

$$K_W = 1 \times 10^{-14}$$

In reality, the equilibrium constant for water, K_W, is an experimentally determined value. The fact that, in pure water, $[H^+_{(aq)}] = [OH^-_{(aq)}] = 1 \times 10^{-7}$ mol/L follows as a result of this experimentally determined value.

SAMPLE QUESTIONS

1. *What is the $[H^+]$ and $[OH^-]$ of pure water at equilibrium?*

2. *Contrary to what one might think, pure water does not contain only $H_2O_{(l)}$ molecules. Actually at 25°C, for every litre of water, there are 1.0 x 10^{-7} moles of water molecules which dissociate according to the equation:*
   $H_2O_{(l)} \leftrightarrow H^+(aq) + OH^-(aq)$

   Calculate the K_W for this system.

3. *Black coffee has a hydronium ion concentration about 100 times greater than pure water.*

   Which of the following is a possible pH for black coffee?

   A) 2.3 B) 4.9 C) 8.9 D) 10.2

4. *According to Le Châtelier's Principle, what would be the effect of adding 10 mL of a strong acid to 990 mL of pure water?*

   A) The product of the $[H^+_{(aq)}]$ and $[OH^-(aq)]$ in the resulting solution would be greater than the product of the $[H^+_{(aq)}]$ and the $[OH^-(aq)]$ in pure water.
   B) The $[H^+_{(aq)}]$ in the resulting solution would be greater than the $[H^+_{(aq)}]$ in pure water.
   C) The $[H^+_{(aq)}]$ in the resulting solution would be greater than the $[H^+_{(aq)}]$ in the original acid solution.
   D) The $[OH^-_{(aq)}]$ in the resulting solution would be greater than the $[OH^-_{(aq)}]$ in pure water.

2.5 To compare the strength of various acids, based on experiments and using simulations.

KEY CONCEPTS:

ACID/BASE REVIEW

Properties of Acids:

1. Conduct electricity
2. React with some metals to release hydrogen
3. Turn blue litmus paper to red
4. Taste sour
5. React with a base to produce a salt and water (Neutralization)

$$\underset{\text{acid}}{HCl(aq)} + \underset{\text{base}}{NaOH(aq)} \rightarrow \underset{\text{salt}}{NaCl(aq)} + \underset{\text{water}}{H_2O(l)}$$

6. Ionize in solution to liberate hydrogen ions, $H^+(aq)$

$$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

$$H_2CO_{3(aq)} \rightarrow 2H^+_{(aq)} + CO_3^{2-}(aq)$$

Examples of acids: H_2SO_4, HNO_3, H_3PO_4, H_2CO_3, HCl, CH_3COOH ($HC_2H_3O_2$)

Properties of Bases:

1. Conduct electricity
2. Turn phenolphthalein red
3. Turn red litmus paper to blue
4. Taste bitter and feel slippery to touch
5. React with acids to produce a salt and water (Neutralization)

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

Acidbase salt water

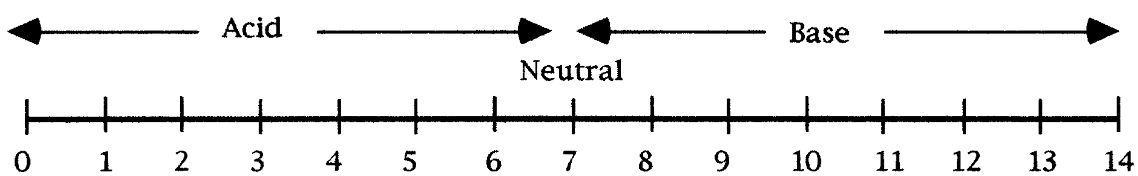
6. Ionize in solution to liberate hydroxide ions, $OH^-(aq)$

$$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

$$NH_4OH_{(aq)} \rightarrow NH_4^+{}_{(aq)} + OH^-_{(aq)}$$

Examples of bases: NaOH, $Ca(OH)_2$, NH_4OH, KOH

pH scale



A pH scale is used to show how acidic or basic a solution is.

- can be used to measure the concentration of hydrogen and hydroxide ions in a solution
- can be found using the mathematical formula:

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

Titration is the volumetric process used to calculate the concentration of an unknown acid or base using a neutralization reaction.

Summary of $[H^+]$, $[OH^-]$ and pH

$[H^+]$	10^0	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}
pH	**0**	**1**	**2**	**3**	**4**	**5**	**6**	**7**
$[OH^-]$	10^{-14}	10^{-13}	10^{-12}	10^{-11}	10^{-10}	10^{-9}	10^{-8}	10^{-7}

$[H^+]$	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}
pH	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**
$[OH^-]$	10^{-7}	10^{-6}	10^{-5}	10^{-4}	10^{-3}	10^{-2}	10^{-1}	10^0

Higher $[H^+]$ more acidic (lower pH)
Higher $[OH^-]$ more basic (higher pH)

Water has a pH = 7 and a pOH = 7, that is:

$[H^+] = 1.0 \times 10^{-7}$ mol/L
$[OH^-] = 1.0 \times 10^{-7}$ mol/L

Therefore $[H^+] \times [OH^-] = 1.0 \times 10^{-14}$

OR $pH + pOH = 14$

To find the pH of any solution knowing the $[H^+]$.

Example 1

Find the pH of a solution with hydrogen ion concentration of 3.0×10^{-4} mol/L.

$pH = -\log(3.0 \times 10^{-4}) = 3.5$

Solution

To use calculator correctly;

$\boxed{-}$ $\boxed{\log}$ $\boxed{3}$ $\boxed{EXP}$ $\boxed{-}$ $\boxed{4}$ $\boxed{=}$

Example 2

Find the pH of a solution with hydroxide ion concentration of 2.0×10^{-6} mol/L.

Solution

$[OH^-] = 2.0 \times 10^{-6}$ mol/L

$[H^+] \times [OH^-] = 1.0 \times 10^{-14}$

$$[H^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-6}\ mol/L} = 5.0 \times 10^{-9}\ mol/L$$

$pH = -\log(5.0 \times 10^{-9}) = 8.3$

To find the $[H^+]$ from pH.

Example 1

Given the pH of a solution is 2.3 find the $[H^+]$.

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

$2.3 = -\log[H^+]$

$[H^+] = 5.0 \times 10^{-3}$ mol/L

Solution:

To use the calculator correctly;

$\boxed{2nd}$ $\boxed{\log}$ $\boxed{-}$ $\boxed{2.3}$ $\boxed{=}$

Example 2

Given the pOH = 5.9, find the $[H^+]$.

> ***Solution:***
>
> $pH + pOH = 14$
> $pH = 14 - 5.9 = 8.1$
>
> $pH = - \log [H^+]$
> $8.1 = - \log [H^+]$
>
> $[H^+] = 7.9 \times 10^{-9}$ mol /L

Equilibrium Constant for an Acid (K_a)

Reflects that fraction of an acid that is dissociated to liberate $[H^+]$.

$$HA_{(aq)} \rightarrow H^+_{(aq)} + A^-_{(aq)}$$

$$K_a = \frac{[H^+_{(aq)}][A^-_{(aq)}]}{[HA_{(aq)}]}$$

If the K_a has a small value, the dissociation of the acid molecules is low. Relatively few hydrogen ions are present at equilibrium. This is a weak acid.

If the K_a has a large value, the dissociation of the acid molecules is high. A relatively large number of hydrogen ions are present at equilibrium. This is a strong acid.

It is important that we do not confuse the concentration of an acid with its strength. The concentration refers to the number of acid molecules or the ions that they produce that are present in a given volume of solution. The strength of an acid refers to the extent to which the acid molecules dissociate to form hydrogen ions.

The concentration of hydrogen ions in a solution depends on both the concentration and the strength of the acid. Therefore, is it possible to have concentrated weak acid with a larger hydrogen ion concentration and a lower pH than a diluted strong acid.

Example 1

A 0.100 mol/L solution of acetic acid is partially ionized. At equilibrium, the $[H^+] = 1.34 \times 10^{-3}$ mol/L.

Calculate the equilibrium constant of acetic acid.

Solution:

$$CH_3COOH_{(aq)} \rightarrow CH_3COO^-_{(aq)} + H^+_{(aq)}$$

	$CH_3COOH_{(aq)}$ →	$CH_3COO^-_{(aq)}$ +	$H^+_{(aq)}$
start	0.100 mol/L	0	0
shift	-x	+ x	+ x
equilibrium	0.100 - x 0.100 - 1.34×10^{-3} 9.87×10^{-2} mol/L	1.34×10^{-3} mol/L	1.34×10^{-3} mol/L

$$K_a = \frac{[CH_3COO^-_{(aq)}]\,[H^+_{(aq)}]}{[CH_3COOH_{(aq)}]} = \frac{[1.34 \times 10^{-3}]\,[1.34 \times 10^{-3}]}{[9.87 \times 10^{-2}\ \text{mol/L}]} = 1.82 \times 10^{-5}$$

Example 2

A 0.15 mol/L solution of $HNO_{3(aq)}$ is ionized. At equilibrium, the pH of the solution is 3.

What is the K_a of this solution?

Solution:

$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$

$pH = -\log [H^+]$
$3.0 = -\log [H^+]$
$[H^+] = 1.0 \times 10^{-3}$ mol/L

	$HNO_3(aq)$	$\rightarrow$ $H^+(aq)$	+ $NO_3^-(aq)$
Start	0.15 mol/L	0	0
Shift	-x	+ x	+ x
Equilibrium	0.15 - x $0.15 - 1.0 \times 10^{-3}$ 0.149 mol/L	1.0×10^{-3} mol/L	1.0×10^{-3} mol/L

$$K_a = \frac{[NO_3^-(aq)]\,[H^+(aq)]}{[HNO_3(aq)]} = \frac{[1.0 \times 10^{-3}]\,[1.0 \times 10^{-3}]}{[0.149]} = 6.7 \times 10^{-6}$$

The K_a can be also used to compare relative strength of acids.

Example

A strong acid is a substance which, in an aqueous solution, highly dissociates to produce $H^+(aq)$ ions.

Which of the following is the strongest acid?

Solution:

$HClO_2$ $K_a = 1.2 \times 10^{-2}$
$HClCH_2CO_2$ $K_a = 1.4 \times 10^{-3}$
$HClO$ $K_a = 1.0 \times 10^{-8}$

The strong acid would be $HClO_2$ because it has the largest K_a value, this would mean it highly ionizes in solution.

SAMPLE QUESTIONS

1. *What is the hydroxide ion concentration of a solution with a pH of:*

   a) *3?*
   b) *6?*
   c) *12?*

2. *In an aqueous solution in which the* $[H^+_{(aq)}] = 1.0 \times 10^{-3}$ *mol/L, what will be the* $[OH^-_{(aq)}]$*?*

3. *Arrange the following in increasing order of strength:*

$HF_{(aq)}$	$\leftrightarrow$	$H^+_{(aq)} + F^-_{(aq)}$	$K_a = 3.5 \times 10^{-4}$
$HNO_{2(aq)}$	$\leftrightarrow$	$H^+_{(aq)} + NO_2^-{}_{(aq)}$	$K_a = 4.6 \times 10^{-4}$
$NH_4^+{}_{(aq)}$	$\leftrightarrow$	$H^+_{(aq)} + NH_{3(aq)}$	$K_a = 5.6 \times 10^{-10}$
$HCO_3^-{}_{(aq)}$	$\leftrightarrow$	$H^+_{(aq)} + C0_3^{-2}{}_{(aq)}$	$K_a = 5.6 \times 10^{-11}$

4. *Given the following data collected in the laboratory:*

$HClO_{3(aq)}$	$K_a = 1.2 \times 10^{-2}$
$HClO_{(aq)}$	$K_a = 1.0 \times 10^{-8}$
$HClCH_2CO_{2(aq)}$	$K_a = 1.4 \times 10^{-3}$

   Classify these acids in increasing order of relative strength and justify your answer.

5. *The pH of a 0.100 mol/L solution of* $HF_{(aq)}$ *is 5.50. The ionization of this acid is represented by the following equation:*

   $$HF_{(aq)} \leftrightarrow H^+_{(aq)} + F^-_{(aq)}$$

   Find the ionization constant (K_a)*, for this acid.*

6. *The K_a of HF acid is 6.7×10^{-4} at room temperature. What would be the $H^+_{(aq)}$ ion concentration in a solution of this acid whose initial concentration is 2.0mol/L?*

$$HF_{(aq)} \leftrightarrow H^+_{(aq)} + F^-_{(aq)}$$

A) 1.7×10^{-4} **mol/L**
B) 1.3×10^{-3} **mol/L**
C) 3.4×10^{-4} **mol/L**
D) 3.6×10^{-2} **mol/L**

7. *A 1.00 L volumetric flask contains 600 mL of distilled water to which a student adds 0.40 g of sodium hydroxide, $NaOH_{(s)}$. Once the $NaOH_{(s)}$ has dissolved, he adds distilled water until the flask is filled, keeping the temperature constant. He then seals the flask. The ionization constant for water is 1.0×10^{-14} at the same temperature. What is the pH of the resulting solution?*

8. *A solution of 0.25 mol/L of hydrofluoric acid is found to contain $[H^+_{(aq)}] = 4.0 \times 10^{-3}$ mol/L at equilibrium. Calculate the K_a of this acid.*

$$HF_{(aq)} \leftrightarrow H^+_{(aq)} + F^-_{(aq)}$$

9. *A solution of $NH_4OH_{(aq)}$ has a concentration of 0.12 mol/L and a pH of 10.*

What is the equilibrium constant for this reaction?

$$NH_4OH_{(aq)} \leftrightarrow NH_4^+{}_{(aq)} + OH^-_{(aq)}$$

10. *A student adds some hydrochloric acid, $HCl_{(aq)}$, to water and finds the resulting solution to have a pH of 3.50. What is the hydroxide ion concentration $[OH^-_{(aq)}]$, of this solution? ($K_w = 1.00 \times 10^{-14}$)*

11. Three acids are examined and the following data is collected:

Acid	Equilibrium conc. of acid (mol/L)	pH	K_a
HOCl	0.30	4	
$HC0_2H$		2	1.8×10^{-4}
HOBr	5.0×10^{-2}		2.0×10^{-9}

The dissociation equations for these acids are:

$$HOCl_{(aq)} \leftrightarrow H^+(aq) + OCl^-(aq)$$
$$HCO_2H_{(aq)} \leftrightarrow H^+(aq) + CO_2H^-(aq)$$
$$HOBr(aq) \leftrightarrow H^+(aq) + OBr^-(aq)$$

Calculate the values of the blank spaces in the data chart.

3.1 To describe, based on the concept of energy, the behaviour of substances in a given oxidation-reduction reaction.

KEY CONCEPTS

Electrochemistry deals with conversions between chemical energy and electrical energy. This occurs only in certain reactions. These reactions are usually spontaneous (no external energy is required) and are referred to as **OXIDATION-REDUCTION REACTIONS** (REDOX).

Oxidation Half-reaction

Oxidation: **A** reaction in which electrons are lost

the substance which undergoes oxidation (donates e^-) is called the reducing agent.

$$\underset{\text{reducing agent}}{Mg} \rightarrow Mg^{2+} + 2e^- \quad \text{(loss of } 2e^-\text{)}$$

Reduction Half-reaction

Reduction: **A** reaction in which electrons are gained

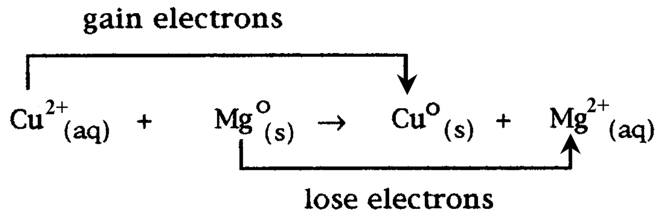
the substance which undergoes reduction (accepts e^-) is called the oxidizing agent.

$$\underset{\text{oxidizing agent}}{S} + 2e^- \rightarrow S^{2-} \quad \text{(gain of } 2e^-\text{)}$$

Note: Both reactions occur at the same time.

In a redox reaction:

$$Cu^{2+}(aq) \quad + \quad Mg(s) \quad \rightarrow \quad Cu(s) \quad + \quad Mg^{2+}(aq)$$



Half-reactions: Represent that part of the reaction which is oxidation and that part of the reaction which is reduction. The electrons are written in to show how they are transferred in the reaction.

$Mg(s) \quad \rightarrow \quad Mg^{2+}(aq) \quad + \quad 2e^-$ oxidation half-reaction

$Cu^{2+}(aq) \quad + \quad 2e^- \quad \rightarrow \quad Cu(s)$ reduction half-reaction

The electrons appear on the right in an oxidation half-reaction.
The species which is losing the electrons becomes more positive (or less negative).
The oxidation number increases.
The substance which is losing the electrons is the reducing agent.

The electrons appear on the left in a reduction half-reaction.
The species which is gaining the electrons becomes less positive (or more negative).
The oxidation number decreases.
The substance which is gaining the electrons is the oxidizing agent.

Example 1

$2AgNO_{3(aq)} + Cu_{(s)} \rightarrow Cu(NO_3)_{2(aq)} + 2Ag_{(s)}$

$2Ag^{+}(aq) + \cancel{2NO_3^{-}}(aq) + Cu^{\circ}(s) \rightarrow Cu^{2+}(aq) + \cancel{2NO_3^{-}}(aq) + 2Ag^{\circ}(s)$

$NO_3^{-}{}_{(aq)}$ are spectator ions since there is no exchange of electrons (their oxidation number does not change).

Redox reaction:

$2Ag^{+}(aq) \quad + \quad Cu^{\circ}(s) \quad \rightarrow \quad Cu^{2+}(aq) \quad + \quad 2Ag^{\circ}(s)$

gain electrons

$$2Ag^{+}_{(aq)} + Cu^{\circ}_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2Ag^{\circ}_{(s)}$$

lose electrons

$Cu^{\circ}(s) \quad \rightarrow \quad Cu^{2+}(aq) \quad + \quad 2e^{-}$ oxidation reaction
reducing agent

$2Ag^{+}(aq) \quad + \quad 2e^{-} \quad \rightarrow \quad 2Ag^{\circ}(s)$ reduction reaction
oxidizing agent

Example 2

$PbCl_{2(aq)} + K_2SO_{4(aq)} \rightarrow 2KCl(aq) + PbSO_{4(s)}$

$Pb^{+2}{}_{(aq)} + 2Cl^{-}{}_{(aq)} + 2K^{+}{}_{(aq)} + SO_4^{-2}{}_{(aq)} \rightarrow 2K^{+}{}_{(aq)} + 2Cl^{-}{}_{(aq)} + PbSO_{4(s)}$

None of the elements have changed oxidation number, therefore this reaction is <u>not a redox reaction</u>.

SAMPLE QUESTIONS

1. *Given:* $Al_{(s)} + Fe^{3+}{}_{(aq)} \rightarrow Fe^{2+}{}_{(aq)} + Al^{3+}{}_{(aq)}$

   Write the reduction equation.
   Write the oxidation equation.
   Name the reducing agent.
   Name the oxidizing agent.

2. *Given:* $Mg_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Mg^{2+}{}_{(aq)} + Cu_{(s)}$

   Write the reduction equation.
   Write the oxidation equation.
   Name the reducing agent.
   Name the oxidizing agent.

3. *Given:* $Au^{3+}{}_{(aq)} + Cd_{(s)} \rightarrow Au_{(s)} + Cd^{2+}{}_{(aq)}$

   Write the reduction equation.
   Write the oxidation equation.
   Name the reducing agent.
   Name the oxidizing agent.

4. *Which of the following statements concerning electrochemical cells is FALSE?*

   A) Electrodes are conductors at which oxidation and reduction occur.
   B) Oxidation takes place at the anode.
   C) Reduction is a partial reaction involving a gain of electrons.
   D) The reducing agent undergoes reduction.

5. *Given the following oxidation-reduction reaction:*

   $Zn_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Zn^{2+}{}_{(aq)} + Cu_{(s)}$

   Which of the following statements is true?

   1) *The zinc is the oxidizing agent.*
   2) *The Cu^{+2} ions are the reducing agent.*
   3) *Zinc atoms are oxidized.*
   4) *Cu^{+2} ions are oxidized.*
   5) *Each zinc atom loses $2e^{-}$ in the reaction.*

   A) 1 and 3 B) 1 and 5 C) 2 and 4 D) 3 and 5

6. *Which of the following definitions is/are TRUE?*

   A) The oxidizing agent causes reduction.
   B) A negative ion is called a cation.
   C) The anode is the electrode where reduction takes place.
   D) The half-reaction in which there is a gain of electrons is called reduction.

7. *Magnesium reacts with $Zn^{+2}{}_{(aq)}$ ions but not with $Sr^{+2}{}_{(aq)}$ ions. Aluminum reacts with $Zn^{+2}{}_{(aq)}$ ions but not with $Mg^{+2}{}_{(aq)}$ ions.*

   Arrange these elements in decreasing order of their tendency to undergo oxidation.

   A) Sr, Mg, Al, Zn
   B) Mg, Sr, Zn, Al
   C) Zn, Al, Mg, Sr
   D) Al, Zn, Sr, Mg

8. *You conduct tests in the laboratory to verify whether the metals aluminum, $Al_{(s)}$, copper, $Cu_{(s)}$, nickel, $Ni_{(s)}$ and zinc, $Zn_{(s)}$, react on contact with solutions of $Al(NO_3)_{3(aq)}$, $Cu(NO_3)_{2(aq)}$, $Zn(NO_3)_{2(aq)}$, and $Ni(NO_3)_{2(aq)}$. The table below summarizes your observations.*

Metal Solution	$Al_{(s)}$	$Cu_{(s)}$	$Ni_{(s)}$	$Zn_{(s)}$
$Al(NO_3)_{3(aq)}$	no	no	no	no
$Cu(NO_3)_{2(aq)}$	yes	no	yes	yes
$Zn(NO_3)_{2(aq)}$	yes	no	no	no
$Ni(NO_3)_{2(aq)}$	yes	no	no	yes

*Based on the above table, how would you arrange these metallic ions if you put them in **increasing** order of their ability to undergo reduction?*

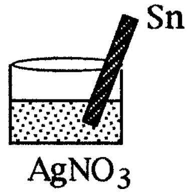
A) Cu^{2+} ***(aq),*** Ni^{2+} ***(aq),*** Zn^{2+} ***(aq) and*** Al^{3+} ***(aq)***
B) Al^{3+} ***(aq),*** Cu^{2+} ***(aq),*** Ni^{2+} ***(aq) and*** Zn^{2+} ***(aq)***
C) Zn^{2+} ***(aq),*** Ni^{2+} ***(aq),*** Cu^{2+} ***(aq) and*** Al^{3+} ***(aq)***
D) Al^{3+} ***(aq),*** Zn^{2+} ***(aq),*** Ni^{2+} ***(aq) and*** Cu^{2+} ***(aq)***

9. *Given the following half-reactions:*

$$Ca^{2+}{}_{(aq)} + 2e^- \rightarrow Ca_{(s)}$$
$$Cr_{(s)} \rightarrow Cr^{3+}{}_{(aq)} + 3e^-$$
$$Sn^{3+}{}_{(aq)} + 1e^- \rightarrow Sn^{2+}{}_{(aq)}$$
$$2Cl^-{}_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$$

Circle the half-reaction(s) which show(s) oxidation.

10. In an experiment, a student prepares a solution of silver nitrate ($AgNO_3$) and puts a thin piece of tin (Sn) into it.



The half-reactions are represented by the following equations:

1. $Sn_{(s)} \rightarrow Sn^{2+}{}_{(aq)} + 2e^-$
2. $Ag_{(aq)} + 1e^- \rightarrow Ag_{(s)}$

Which of the following statements is correct?

A) Equation 1 represents the oxidation of Sn and Sn is the oxidizing agent.
B) Equation 1 represents the reduction of Sn and Sn is the reducing agent.
C) Equation 2 represents the reduction of Ag^+ and Ag^+ is the oxidizing agent.
D) Equation 2 represents the oxidation of Ag^+ and Ag^+ is the reducing agent.

11. A chemical reaction is represented by the following reaction:

$$5Zn_{(s)} + 2NO_3^-{}_{(aq)} + 12H^+{}_{(aq)} \leftrightarrow 5Zn^{+2}{}_{(aq)} + N_{2(g)} + 6H_20_{(l)}$$

Which of the following statements concerning the reaction are TRUE?

1. *The zinc is oxidized.*
2. *The zinc is absorbed.*
3. *The nitrate ion, $NO_3^-{}_{(aq)}$, is the oxidizing agent.*
4. *The $H^+{}_{(aq)}$ ion is the reducing agent.*

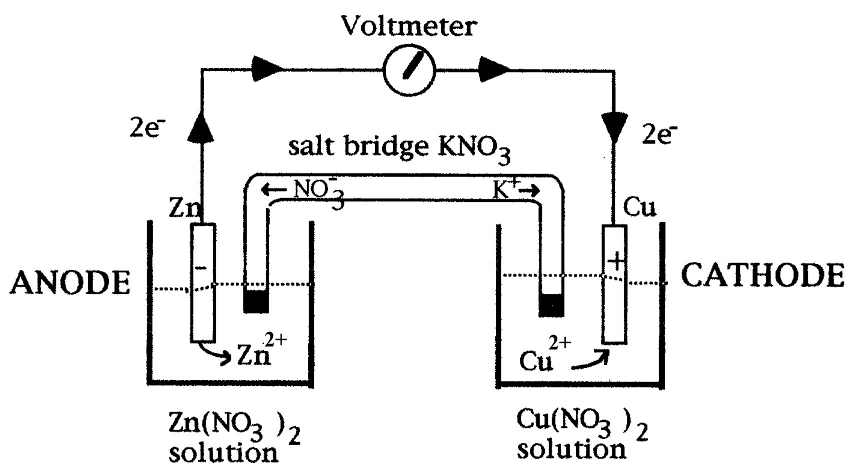
A) 1 and 3 ***B) 1 and 4*** ***C) 2 and 3*** ***D) 2 and 4***

3.2 To determine the pair of electrodes that offer the greatest electrical potential, based on experiments.

KEY CONCEPTS

Electrochemical Cells

If the oxidation process causes electrons to be given off and the reduction process attracts electrons, then by connecting an external wire we can have an electron flow (electric current) through the wire. This can be accomplished by a redox reaction in a device that converts chemical energy into electrical energy spontaneously. This device is called a **VOLTAIC CELL**.



Anode
- electrode where oxidation occurs
- where electrons are lost

Cathode
- electrode where reduction occurs
- where species in solution gain electrons

The flow of electrons in the wire (as indicated by the arrows) is from the anode to the cathode. We usually think of an electric circuit as a flow of electrons from negative to positive. Therefore, the anode of an electrochemical cell is sometimes labeled negative and the cathode is sometimes labeled positive. However, labeling the anode and cathode as negative and positive seems to contradict what was taught in Science 416. The negative electrode in a cathode ray tube was called the cathode. In the electrolysis of water, the negative electrode was also called the cathode. Therefore, when labeling an electrode as the anode or cathode, consider the type of reaction that is taking place. Think of a red cat. (<u>red</u>uction at the <u>cat</u>hode)

Function of Voltaic Cell

1. Electrons are produced at Zn electrode (ANODE)

   $Zn°_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2\ e^-$ oxidation reaction
   reducing agent

   Zn^{2+} ions are removed into the solution. The Zn electrode loses mass.

2. Electrons travel through the wire and the voltmeter measures the voltage (the energy that each unit of charge can release). Because the electrons can release energy, the current can do work (make a light glow, power a radio, etc.)

3. Electrons enter Cu electrode (CATHODE) and picked up by the Cu^{2+} ions in the solution.

   $Cu^{2+}{}_{(aq)} + 2\ e^- \rightarrow Cu°_{(s)}$ reduction reaction

   Cu^{2+} ions accept electrons and plate on the Cu electrode.
   The Cu electrode gains mass.

4. To complete the circuit, ions (positive and negative) move freely through an aqueous solution via the salt bridge. The salt bridge contains a conducting solution which allows the passage of ions from one compartment to another but prevents the solutions from mixing.

   NET OXIDATION-REDUCTION REACTION

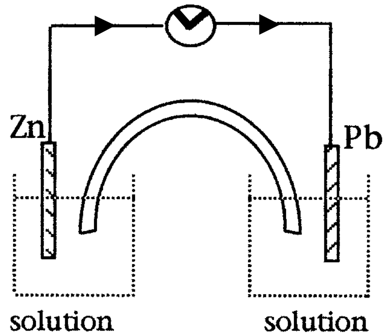
   $Cu^{2+}{}_{(aq)} + Zn_{(s)} \rightarrow Cu_{(s)} + Zn^{2+}{}_{(aq)}$

   OR $Zn_{(s)}\ |\ Zn(NO_3)_{2(aq)}\ \|\ Cu(NO_3)_{2(aq)}\ |\ Cu_{(s)}$

   OR $Zn_{(s)}\ |\ Zn^{2+}{}_{(aq)}\ \|\ Cu^{2+}{}_{(aq)}\ |\ Cu_{(s)}$

Example

Given:

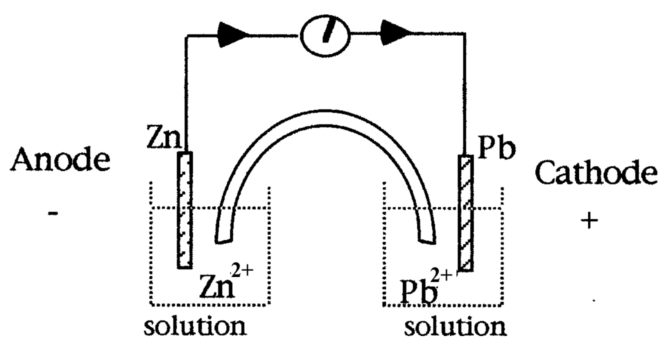


Find:

Anode, cathode, electrode where oxidation occurs, electrode where reduction occurs, half-reactions, oxidizing agent and reducing agent.

Solution:

Because the arrows point to the right the flow of electrons is from the Zn electrode to the Pb electrode, therefore;



$Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2e^-$ oxidation

reducing agent

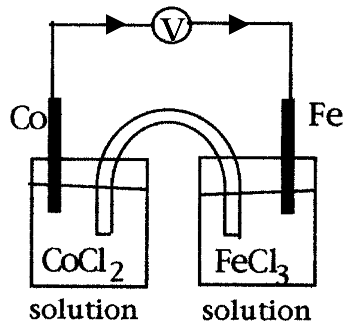
$Pb^{2+}{}_{(aq)} + 2e^- \rightarrow Pb_{(s)}$ reduction

oxidizing agent

NET EQUATION $Pb^{2+}{}_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + Pb_{(s)}$

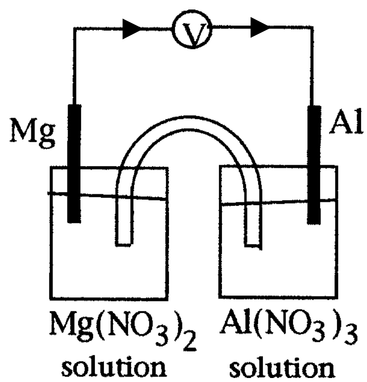
SAMPLE QUESTIONS

1. *Given a Co | Co^{2+} || Fe^{3+} | Fe voltaic cell:*



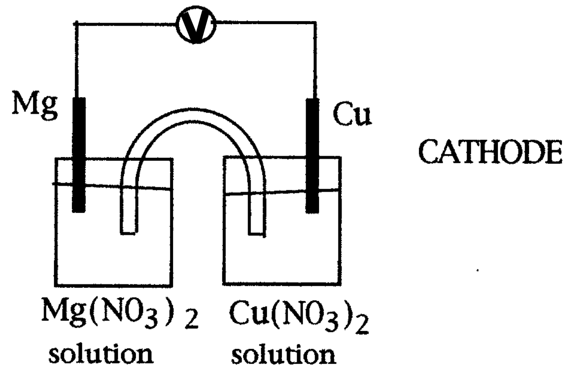
   a) *Indicate the anode.*
   b) *Write the half-reactions and the net equation for this cell.*

2. *The following diagram illustrates a voltaic cell.*

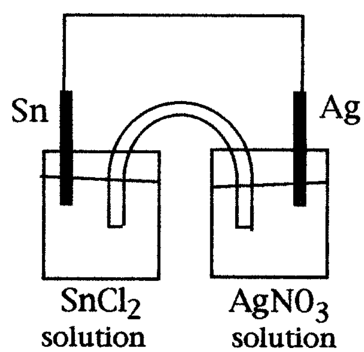


Write out the half cell reaction that occurs at the cathode.
Write out the complete balanced equation for the net equation.
Identify the electrode which increases in mass.

3. *Given the following voltaic cell indicate the direction of the electron flow. Note: Cu is the cathode.*



4. *The diagram below represents an electrochemical cell.*



$$Sn_{(s)} \rightarrow Sn^{2+}{}_{(aq)} + 2e^-$$
$$Ag^+{}_{(aq)} + 1e^- \rightarrow Ag_{(s)}$$

Which of the following statements is FALSE?

A) Electrons move through the wire from the Sn electrode to the Ag electrode.
B) The Sn electrode is the cathode.
C) Silver forms on the Ag electrode.
D) The $NO_3^-{}_{(aq)}$ ion moves through the salt bridge towards the Sn electrode in the solution.

3.3 Calculate the potential difference of various oxidation-reduction reaction, based on a table of standard electrode potentials.

KEY CONCEPTS

Electric Potential

Electric Potential (voltage) is a measure of the energy given to each unit of charge. One volt of electric potential is equal to one joule of energy per coulomb of charge. When electrons are given a lot of energy, they have a tendency to flow to an area where the energy can be released. Therefore, the electric potential of a voltaic cell relates to its ability to produce an electric current.

Metals can be classified by their ability to donate electrons (oxidation) and/or their ability to accept electrons (reduction). Chemists have agreed to compare metals by their ability to gain electrons (<u>reduction potential</u>) as compared to the hydrogen half-reaction.

$H_{2(g)} \rightarrow 2H^+_{(aq)} + 2e^-$ Electric potential 0 V or $E° = 0$ Volts

In a voltaic cell we can obtain differences in potentials between two metals as compared to the reduction potential of the half-reaction of hydrogen, then compare one metal with another and decide which will be reduced and which oxidized. A table of reduction potentials is used to obtain these values.

Example

Given the following voltaic cell: $Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$,
state the oxidation half-reaction and the reduction half-reaction and indicate the correct E° value for each half-reaction. (Refer to the table of Reduction Potentials.)

Solution:

From reduction potential table.

$Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ $E° = +0.34$ V reduction

$Zn^{2+}{}_{(aq)} + 2e^- \rightarrow Zn_{(s)}$ $E° = -0.76$ V oxidation

The metal with the higher reduction potential will be reduced and the metal with the lower reduction potential will be oxidized. This means that copper has a greater tendency to accept electrons than zinc, therefore zinc will donate its electrons to copper. Zinc undergoes an oxidation reaction so its half-equation is rewritten to show oxidation and the sign of its electric potential is changed to reflect this.

$Cu^{2+}{}_{(aq)} \quad + \quad 2e^- \quad \rightarrow \quad Cu_{(s)}$ $\qquad$ E° = +0.34 V

$Zn_{(s)} \quad \rightarrow \quad Zn^{2+}{}_{(aq)} \quad + \quad 2e^-$ $\qquad$ E° = +0.76 V (change sign)

Calculating Standard Cell Potential

The sum of the two half-reactions which results in a REDOX reaction can be used to determine the standard potential for the cell. The standard cell potential is calculated as the sum of the electric potential of the reduction and oxidation half-reactions.

E°cell = E°reduction + E°oxidation

$Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ $\qquad$ E° = +0.34 V

$Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2e^-$ $\qquad$ E° = +0.76 V

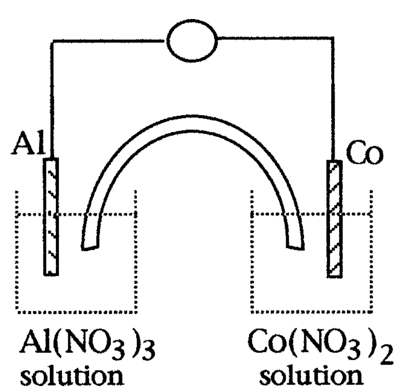
NET: $Cu^{2+}{}_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + Cu_{(s)}$ $\qquad$ E° = + 1.10 V

If the standard cell potential value is positive the reaction is spontaneous.
If the standard potential value is negative the reaction is not spontaneous and will not occur.

Example

Given the voltaic cell:

$Al_{(s)} \mid Al(NO_3)_{3(aq)} \parallel Co(NO_3)_{2(aq)} \mid Co_{(s)}$



Find the standard cell potential for this cell and state the oxidizing agent and the reducing agent.

Solution:

From the reduction potential table;

$Al^{3+}_{(aq)} + 3e^- \rightarrow Al_{(s)}$ E° = -1.66 V oxidation

$Co^{2+}_{(aq)} + 2e^- \rightarrow Co_{(s)}$ E° = -0.28 V reduction

Rewriting equations and balancing the half-reactions

$3(Co^{2+}_{(aq)} + 2e^- \rightarrow Co_{(s)})$ E° = -0.28 V

$2(Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^-)$ E° = +1.66 V **(change sign)**

NET: $3Co^{2+}_{(aq)} + 2Al_{(s)} \rightarrow 2Al^{3+}_{(aq)} + 3Co_{(s)}$ E° = +1.38 V
reaction is spontaneous

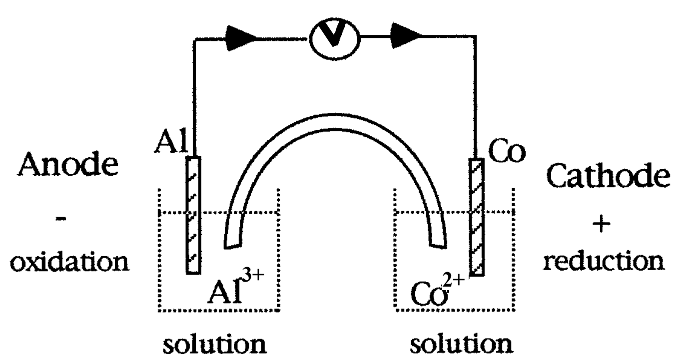
$Co^{2+}_{(aq)}$ is the oxidizing agent

$Al_{(s)}$ is the reducing agent

Note: When balancing the half-reactions, the standard potential of the reaction is not multiplied.

When a voltaic cell is constructed and the connections are made, the competition for electrons between the two metals (electron donors) begins.

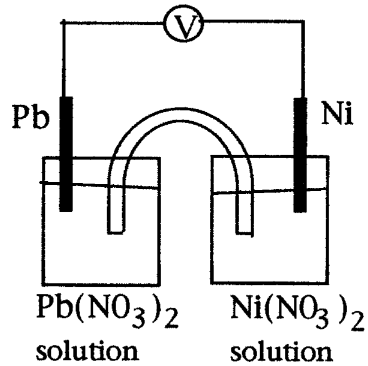
The strong electron donor (lower reduction potential) will donate electrons (oxidation reaction).
The weak electron donor (higher reduction potential) will gain electrons (reduction reaction).



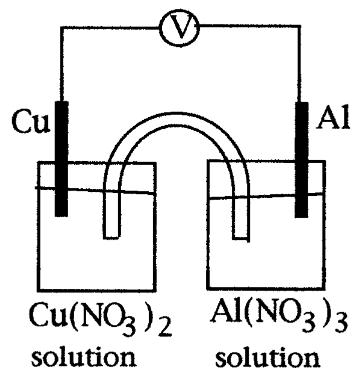
SAMPLE QUESTIONS

For the following questions, refer to reduction potential table.

1. *Given the following voltaic cell indicate which electrode is the* ***ANODE*** *and which is the* ***CATHODE****.*

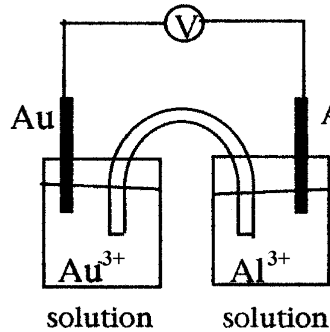


2. *You are presented with a standard cell $Al|Al^{3+} || Cu^{+2}|Cu$. The solutions have a concentration of 1 mol/L and are maintained at 25°C. Calculate the potential difference that the voltmeter would indicate?*



3. *Calculate the standard cell potential of the following voltaic cell using the standard reduction table.*

E^o = ________



4. *Given the following cell fill-in the required information.*

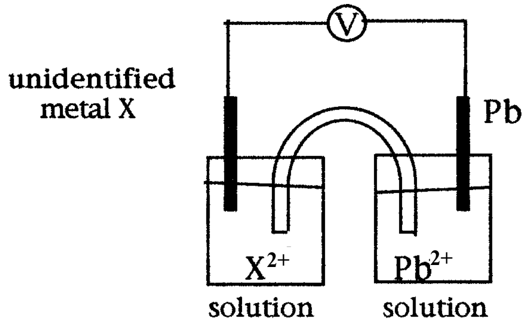
   $Cr_{(s)}$ | $Cr(NO_3)_{2(aq)}$ || $Cu(NO_3)_{2(aq)}$ | $Cu_{(s)}$

   Oxidation reaction:
   Reduction reaction:
   Net reaction:
   The electrode which increases in mass:
   The electrode which decreases in mass:
   The cell voltage:

5. *Calculate the cell potential of these redox reactions and state if reaction will occur spontaneously.*

   a) $Co^{2+}{}_{(aq)} + Fe_{(s)} \rightarrow Fe^{2+}{}_{(aq)} + Co_{(s)}$
   b) $Cu_{(s)} + 2H^{+}{}_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + H_{2(g)}$
   c) $2Ag_{(s)} + Fe^{2+}{}_{(aq)} \rightarrow 2Ag^{+}{}_{(aq)} + Fe_{(s)}$
   d) $3Zn^{2+}{}_{(aq)} + 2Cr_{(s)} \rightarrow 3Zn_{(s)} + 2Cr^{3+}{}_{(aq)}$

6. *Using the diagram below*



If the standard cell voltage of this cell is measured as +0.15 volts, what is the unidentified metal at the anode?

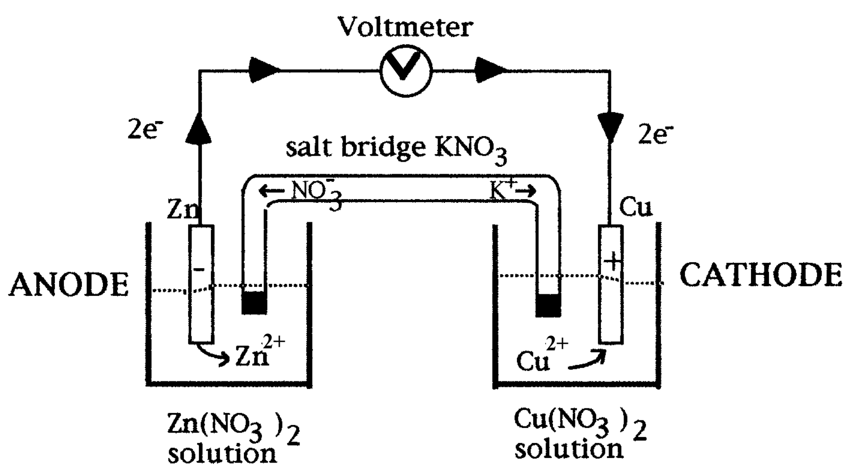
3.4 To explain the variation in the oxidation-reduction potential of an electrochemical cell, based on Le Châtelier's principle.

KEY CONCEPTS

A change in the concentration of the solutions affects cell potential in that it changes the number of ions in the solution, which in turn affects the cell potential. This effect can be explained by Le Châtelier's principle which states if a change occurs in a system at equilibrium, the system will shift in such a way as to counteract this change. In terms of the battery or voltaic cell this leads to a change in electric potential.

Effects of Concentration

Example $Cu^{2+}{}_{(aq)} + Zn_{(s)} \rightarrow Cu_{(s)} + Zn^{2+}{}_{(aq)}$



If the concentration of $Cu(NO_3)_{2(aq)}$, that is the $Cu^{2+}{}_{(aq)}$ ions in the solution increases, the system will shift to remove them from the solution by decomposing (oxidizing) more Zn. This causes more Zn to be oxidized (lose electrons) thus increasing the flow of electrons to the Cu electrode and the electric potential (voltage) goes up. The system shifts to the right.

Increase in concentration of the ions on the reactants side will cause an increase in the cell potential.

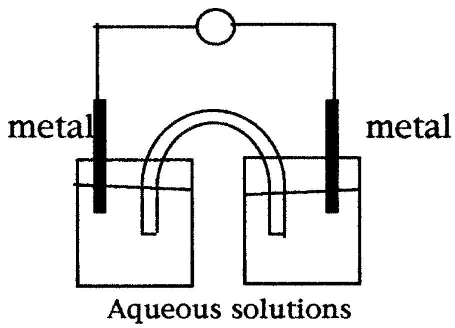
If the concentration of $Zn(NO_3)_{2(aq)}$, that is the $Zn^{2+}{}_{(aq)}$ ions in the solution increases, the system will shift to slow down the oxidation of Zn. This causes less Zn to be oxidized (lose electrons) thus decreasing the flow of electrons to Cu and the electric potential (voltage) goes down. Eventually the flow of electrons will cease and the reaction will stop. This marks the death of the battery or voltaic cell.

Increase in concentration of the ions on the product side will cause a decrease in the cell potential.

SAMPLE QUESTIONS

1. *You wish to use an electrochemical cell to light up a light bulb. The bulb requires a minimum of 1.5 V. The materials listed on this table and shown on the diagram are available.*

Metals	Standard Solutions
Zn	$Zn(NO_3)_2$
Ni	$Ni(NO_3)_2$
Cu	$Cu(NO_3)_2$
Ag	$AgNO_3$



The reduction reactions and potentials appear in the following table.

Reduction Half-Reaction	*E° (Volts)*
$Ag^{+1}{}_{(aq)} + 1e^- \rightarrow Ag°_{(s)}$	+0.80
$Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu°_{(s)}$	+0.34
$Ni^{2+}{}_{(aq)} + 2e^- \rightarrow Ni°_{(s)}$	-0.25
$Zn^{2+}{}_{(aq)} + 2e^- \rightarrow Zn°_{(s)}$	-0.76

Use this table to determine which pair of electrodes you would use to make the electrochemical cell. Explain why.

2. *A scientist must store a 1 mol/L solution of $Cr(NO_3)_{3(aq)}$ at room temperature. Can the scientist use a copper container to store this solution? Explain your answer.*

MODULE 2

ANSWER KEY

OBJECTIVES 1.1 & 1.2

1. A

OBJECTIVES 2.1, 2.2 & 2.3

1. V_1 = 38.0 mL; P_1 = 120 kPa; P_2 = 95.0 kPa; V_2 = ?
$P_1V_1 = P_2V_2$
$V_2 = V_1P_1/P_2$
V_2 = (38.0 mL)(120 kPA)/(95.0 kPa)
V_2 = 48.0 mL

2. V_1 = 80.0 mL; P_1 = 102.4 kPa; P_2 = 100.7 kPa; V_2 = ?
$P_1V_1 = P_2V_2$
$V_2 = P_1V_1/P_2$
V_2 = (80.0 mL)(102.4 kPA)/(100.7 kPa)
V_2 = 81.4 mL

3. V_1 = 8.00 L; T_1 = 25.0°C + 273 = 298 K; T_2 = 50.0°C + 273 = 323 K; V_2 = ?
$V_1/T_1 = V_2/T_2$
$V_2 = V_1T_2/T_1$
V_2 = (8.00 L)(323 K)/(298 K)
V_2 = 8.67 L

4. V_1 = 12.0 mL; T_1 = 115°C + 273 = 388 K; V_2 = 9.00 mL; T_2 = ?
$V_1/T_1 = V_2/T_2$
$T_2 = V_2T_1/V_1$
T_2 = (9.00 mL)(388 K)/(12.0 mL)
T_2 = 291 K – 273 = 18°C

5. V_1 = 52.0 mL; T_1 = 20.0°C + 273 = 293 K; T_2 = 28.0°C + 273 = 301 K; V_2 = ?
$V_1/T_1 = V_2/T_2$
$V_2 = V_1T_2/T_1$
V_2 = (52.0 mL)(301 K)/(293 K)
V_2 = 53.4 mL

OBJECTIVES 2.4, 2.5, 2.6, 2.7 & 2.9

1. A $V = 43.8$ L; $T = 43.0^oC + 273 = 316$ K; $P = 105$ kPa; $R = 8.31$ kPa·L/mol·K
   $PV = nRT$ hence $n = PV/RT$
   $n = (105 \text{ kPa})(43.8 \text{ L})/(8.31 \text{ kPa·L/mol·K})(316 \text{ K})$
   $n = 1.75$ mol

2. C $V_1 = 30.0$ mL; $P_1 = 105$ kPa; $P_2 = 90.0$ kPa
   Since $P_1V_1 = P_2V_2$, it follows that $V_2 = P_1V_1/P_2$
   $V_2 = (30.0 \text{ mL})(105 \text{ kPa})/(90.0 \text{ kPa})$
   $V_2 = 35$ mL or 3.50×10^1 mL

3. C As the pressure exerted by a gas decreases, then the volume increases.

4. B

Mass of empty container + N_2 gas	=	620 g
Mass of empty container	=	480 g
Mass of N_2 gas	=	140 g
Mass of container + unknown gas	=	770 g
Mass of empty container	=	480 g
Mass of unknown gas	=	290 g

   Moles of N_2 = 140 g/ 28 g/mol = 5.0 moles
   Since the temperature and pressure remain constant, there must also be 5.0 moles of the unknown gas present.
   5.0 moles of the unknown gas have a mass of 290 g
   Therefore 1 mole of the unknown gas has a mass of 58 g.
   The molar masses of the suggested answers are:
   C_2H_2 = 26 g/mol; C_4H_{10} = 58 g/mol; C_2H_6 = 30 g/mol; CH_4 = 16 g/mol

5. A

6. B

7. C $V_1 = 4.00 \times 10^2$ mL; $T_1 = -123^oC + 273 = 150$ K; $T_2 = 27.0^oC + 273 = 300$K
   Since $V_1/T_1 = V_2/T_2$, then $V_2 = (4.00 \times 10^2 \text{ mL})(300 \text{ K})/(150 \text{ K})$
   $V_2 = 800$ mL or 8.00×10^2 mL

8. D $V_1 = 4.0$ L; $P_1 = x$; $V_2 = 0.50$ L; find P_2
   Since $P_1V_1 = P_2V_2$, then $P_2 = P_1V_1/P_2$ and $P_2 = (4.0 \text{ L})(x)/0.50 \text{ L})$
   $P_2 = 8x$ or 8 times the original pressure.

9. Mass of syringe + O_2 gas = 80.92 g
Mass of syringe = 80.77 g
Mass of O_2 gas = 0.15 g
Mass of syringe + unknown gas = 81.07 g
Mass of syringe = 80.77 g
Mass of unknown gas = 0.30 g
For oxygen, the following conditions exist:
P = 101.3 kPa; V = 0.113 L; T = 22.0°C + 273 = 295 K; R = 8.31 kPa·L/mol·K
Since PV = nRT ; n = PV/RT
n = (101.3 kPa)(0.113 L)/(8.31 kPa·L/mol·K)(295 K)
n = 4.669 x 10^{-3} mol O_2
Since there are the same conditions of temperature and pressure for both gases, there must also be 4.669 x 10^{-3} mol of the unknown gas present.
Hence if 4.669 x 10^{-3} mol of unknown gas has a mass of 0.30 g then 1 mol of the unknown gas will have a mass of 64 grams.

10. If the volume were reduced to ½ of the original, then the pressure would be double. If at the same time the absolute temperature were tripled, then the pressure would increase by a factor of three. Combining these two ideas indicates that the final pressure would be six times the original.

11. The balanced equation for the reaction is:
$Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$
Since 30.06 g of Calcium are used, the number of moles can be calculated:
30.06 g /40.08 g/mol = 0.75 mol
From the balanced equation, 1 mol of Ca produces 1 mol of H_2. This allows the use of the ideal gas equation: PV = nRT
n = 0.75 mol; R = 8.31 kPa·L/mol·K; T = 22.0°C + 273 = 295 K; P = 100.0 kPa
Since PV = nRT, then V = nRT/P
V = (0.75 mol)(8.31 kPa·L/mol·K)(295 K)/(100.0 kPa)
V = 18.4 L

12. Since the temperature and volume of the two cylinders are identical, the pressure in each cylinder is proportional to the number of moles of gas present.
$P_A/P_B = n_A/n_B$

For nitrogen (A), n_A = 10.0 g/28 g/mol = 0.357 mol
For carbon dioxide (B), n_B = 12 g/44 g/mol = 0.273 mol
Since there are more moles of nitrogen, the pressure in the nitrogen cylinder is greater.

13. From the equation: $C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$
the molar ratio between C_3H_8 and O_2 is 1:5. The molar mass of C_3H_8 is 44.0g/mol. Since there are 801 g of C_3H_8 available, there are 801 g/44.0g/mol or 18.2 mol of C_3H_8. Hence there must be five times this quantity of O_2, or 91.0 mol of O_2 present. Using the ideal gas equation PV = nRT, we can calculate the volume of oxygen required: V = nRT/P
$V = (91.0\ mol)(8.31\ kPa{\cdot}L/mol{\cdot}K)(298\ K)/(101\ kPa)$
$V = 2.23 \times 10^3$ L of oxygen required.

14. If the data collected is used, it is possible to calculate the Universal Gas Constant, R, from the given information. Since PV = nRT, then R = PV/nT
P = 404 kPa; V = 10 L; n = 2.0 mol; T = -73.0°C + 273 = 200 K
Hence R = (404 kPa)(10 L)/(2.0 mol)(200 K)
R = 10.1 kPa·L/mol·K which is an incorrect value for this constant.

15.

Mass of cylinder + oxygen gas	= 148.78 g
Mass of cylinder	= 143.25 g
Mass of oxygen gas	= 5.53 g
Mass cylinder + unknown gas	= 161.91 g
Mass of cylinder	= 143.25 g
Mass of unknown gas	= 18.66 g
Moles of oxygen = 5.53 g/32 g/mol	= 0.173 mol O_2

Since the same temperature, pressure and volume exists, there must be the same number of moles of unknown gas present.
Hence 0.173 mol of the unknown gas has a mass of 18.66 g, then the mass of 1 mol of the unknown gas is 108 grams. There are three suggested gases given in the problem: SO_2 - molar mass 64 g/mol
NO_2 - molar mass 46 g/mol
N_2O_5 - molar mass 108 g/mol corresponding to the answer.

16. Given: T = 23.0°C + 273 = 296 K; V = 0.0473 L; P = 98.4 kPa;
R = 8.31 kPa·L/mol·K, we can calculate the number of moles of hydrogen gas produced by the ideal gas equation: n = PV/RT
$n = (98.4\ kPa)(0.0473\ L)/(8.31\ kPa{\cdot}L/mol{\cdot}K)(296\ K)$
$n = 1.89 \times 10^{-3}$ mol of H_2
From the balanced equation:
$2\ Fe(s) + 3H_2SO_4(aq) \rightarrow Fe_2(SO_4)_3(aq) + 3\ H_2(g)$ hydrogen and iron are in a ratio of 3:2 moles. There must, therefore, be $2/3(1.91 \times 10^{-3}$ mol) of Fe present, that is, 1.26×10^{-3} mol of Fe. Since the molar mass of Fe is 55.85g/mol, this represents 0.0705 g of Fe.

17. The volume and the pressure are constant in this problem, we may use the Universal Gas Equation to solve this problem
$P_1V_1/n_1T_1 = P_2V_2/n_2T_2$. Since both the volume and the pressure do not vary, they may be eliminated from the equation. The equation then simply becomes
$n_1T_1 = n_2T_2$
Ne: 8.40 g/20.18 g/mol = 0.416 mol of Ne
CO_2: 8.40 g/44.01 g/mol = 0.191 mol CO_2
Let n_1 = 0.416 g; T_1 = 22.0°C + 273 = 295 K; n_2 = 0.091 mol. Find T_2.
$T_2 = n_1T_1/n_2$; T_2 = (0.416 mol)(295 K)/(0.191 mol) = 643 K - 273 = 370°C

18. $2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$
Volume of N_2 = 5.60 x 10^4 mL = 56.0 L; T = 25.0°C + 273 = 298 K; P = 101.3 kPa: R = 8.31 kPa·L/mol·K. By the ideal gas equation, n = PV/RT
n = (101.3 kPa)(56.0 L)/(8.31 kPa·L/mol·K)(298 K)
n = 2.29 mol N_2
From the balanced equation N_2 and NaN_3 are in a ratio of 3:2
Hence there will be 2/3(2.29 mol of NaN_3) = 1.53 mol NaN_3
Since the molar mass of NaN_3 is 65.0 g/mol, this value corresponds to 99.5 g of NaN_3.

19. V = 1.50 x 10^3 mL = 1.50 L ; T = 31°C + 273 = 304 K; P = 101.8 kPa;
R = 8.31 kPa·L/mol·K. Use the ideal gas equation to find n.
n = (101.8 kPa)(1.50 L)/(8.31 kPa·L/mol·K)(304 K)
n = 0.0604 mol of air outside.
Since the same amount of air is used inside, there are still 0.0604 mol of air available, i.e. n = 0.0604 mol; V = 2.00 L; P = 72.3 kPa; R = 8.31 kPa·L/mol·K
Solve for T: T = (72.3 kPa)(2.00 L)/(0.0604 mol)(8.31 kPa·l/mol·K)
T = 288 K. This corresponds to 15°C, which represents a drop of 16°C.

20. Mass of syringe + gas = 50.89 g
Mass of syringe = 50.78 g
Mass of gas = 0.11 g
Volume of syringe = 175.0 mL = 0.175 L
Temperature = 21.0°C + 273 = 294 K
Pressure = 96.0 kPa
Use PV = nRT to find the number of moles
n = PV/RT; n = (96.0 kPa)(0.175 L)/(8.31 kPa·L/mol·K)(294 K)
n = 6.88 x 10^{-3} mol of gas. This corresponds to 0.11 grams of the gas. Hence 1 mol of the gas will have a mass of 15.996 g or 16 grams. Calculating the molar masses of the suggested gases, we find the molar mass of methane, CH_4, corresponds to this value.

21. $2HgO(s) \rightarrow 2\,Hg(l) + O_2(g)$
Molar mass of HgO = 216.6 g/mol
There are 8.66 g of HgO given, therefore 8.66g/216.6 g/mol = 0.0400 mol of HgO
From the balanced equation, the ratio of $HgO{:}O_2$ is 2:1, thus there are 0.0200 mol of O_2 generated in the reaction. Using the ideal gas equation, we can calculate the volume of the gas under the given conditions.
$PV = nRT$; $V = nRT/P$ $n = 0.0200$ mol; $R = 8.31$ kPa·L/mol·K;
$T = 30.0^{o}C + 273 = 303$ K; $P = 95.0$ kPa
$V = (0.0200\text{ mol})(8.31\text{ kPa·L/mol·K})(303\text{ K})/(95.0\text{ kPa})$
$V = 0.530$ L

22. Volume and number of moles are constant. Using the Universal Gas Equation, both volume and the number of moles can be eliminated. Hence our equation becomes $P_1/T_1 = P_2/T_2$
$P_1 = 325$ kPa; $T_1 = -12.0^{o}C + 273 = 261$ K; $T_2 = 45.0^{o}C + 273 = 318$ K
$P_2 = P_1T_2/T_1$
$P_2 = (325\text{ kPa})(318\text{ K})/(261\text{ K})$
$P_2 = 396$ kPa

OBJECTIVES 3.1, 3.2 & 3.3

1. A
2. B
3. D

MODULE 3

ANSWER KEY

Review

1.	B	2, 4 and 5 are changes of state (physical changes) and 6 involves the rearrangement of the nuclear particles which results in the formation of isotopes of different elements (nuclear change).
2.	B	

OBJECTIVE 3.1

1.	C	correct by definition.

OBJECTIVES 3.2, 3.3, 1.1, 1.2 and 1.3

1.	D	correct by definition.
2.	D	correct by definition.
3.	C	B and D are endothermic changes of state. In A, the fact that the water cools down means that it has lost heat to the ammonium chloride. The process of dissolving ammonium chloride is therefore endothermic.
4.	B	For 1), burning is <u>always</u> exothermic. For 4), sulfuric acid has released heat to the water (an exothermic process).
5.	D	The process of melting (solid to liquid) in A and C is endothermic. The process of evaporation (liquid to vapour) in B is endothermic.
6.	A	correct by definition.
7.	A	correct by definition.
8.	C	
9.	D	Endothermic reactions are indicated either with a positive ΔH, or with heat shown as a reactant.
10.	B	The energy content of the products is higher than the energy content of the reactants in an endothermic reaction. Thermochemical reactions are always indicated in the direction of reactants yielding products.
11.	B	If the water feels cold to the touch, it has lost heat to the ammonium chloride. The ammonium chloride has absorbed heat: therefore the heat of solution of ammonium chloride is <u>positive</u>, which indicates an <u>endothermic</u> reaction.
12.	B	If the water has increased in temperature, it has absorbed heat from the KOH, which has released heat. The heat of solution of the KOH is negative, which indicates an exothermic reaction. The enthalpy of the

reactants is higher than the enthalpy for the products in an exothermic reaction.

13. D If the enthalpy of the products is lower than that of the reactants, the reaction if exothermic. This is indicated either by a negative ΔH, or with heat shown as a product.

14. a) $\Delta H = \Delta H_{\text{(Products)}} - \Delta H_{\text{(Reactants)}}$

$\Delta H = 50 \text{ kJ/mol} - 300 \text{ kJ/mol}$

$\Delta H = -250 \text{ kJ/mol}$

b) Exothermic

OBJECTIVES 2.1 & 2.6

1. **D** $T_i = 26.0^{o}C$

$T_f = 66.0^{o}C \quad \Delta T = T_f - T_i = 66.0^{o}C - 26.0^{o}C = 40.0\ ^{o}C$

$Q = 36.0$ kJ or 3.60×10^{4} J $\qquad c_{Al} = 0.900\ J/g^{o}C$

$Q = mc\Delta T$

$3.60 \times 10^{4}\ J = m\,(0.900\ J/g^{o}C)(40.0\ ^{o}C)$

$m = 1.00 \times 10^{3}$ g.

2. **D**

water	KOH
$m = 200.0$ g	$m = 4.0$ g
$T_i = 25.0\ ^{o}C$	molar mass = 56.1 g/mol
$T_f = 31.5\ ^{o}C$	
$\Delta T = 6.5\ ^{o}C$	

$Q = mc\Delta T$

$Q = (200.0\ g)(4.19\ J/g^{o}C)(6.5\ ^{o}C)$

$Q = 5450$ J or 5.45 kJ

$$\frac{Q\text{ (water)}}{\Delta m\text{ (KOH)}} = \frac{\Delta H\text{ (KOH)}}{\text{molar mass (KOH)}}$$

$$\frac{5.45\ kJ}{4.0\ g} = \frac{\Delta H\text{ (KOH)}}{56.1\ g/mol}$$

$$\Delta H\text{ (KOH)} = 76\ kJ/mol$$

3. **A**

water

$m = 400.0$ g
$T_i = 20.8\ ^oC$
$T_f = 27.4\ ^oC$
$\Delta T = 6.6\ ^oC$

NaOH and HCl

$n = 0.50$ mol/L x 0.200 L
$n = 0.10$ mol

$Q = mc\Delta T$
$Q = (400.0\ g)(4.19\ J/g^oC)(6.6\ ^oC)$
$Q = 11000\ J = 11 kJ$
$\Delta H = 11kJ/0.10mol = 1.1 \times 10^2$ kJ/mol

4. **C**

$Q = mc\Delta T$
$50400\ J = (800.0\ g)(4.19\ J/g^oC)(T_f - 27.0\ ^oC)$
$T_f - 27.0\ ^oC = 15.0\ ^oC$
$T_f = 42.0\ ^oC$

5. **D**

water

$m = 2000$ g
$T_i = 25.0\ ^oC$
$T_f = 60.7\ ^oC$
$\Delta T = 35.7\ ^oC$

butane

$\Delta m = 1.25g - 7.25$ g
$\Delta m = - 6.00$ g

molar mass = 58.12 g/mol

$Q = mc\Delta T$
$Q = (2000\ g)(4.19\ J/g^oC)(35.7\ ^oC)$
$Q = 299000\ J = 299$ kJ

$$\frac{Q(\text{water})}{\Delta m\ (\text{butane})} = \frac{\Delta H\ (\text{butane})}{\text{molar mass (butane)}}$$

$$\frac{299\ kJ}{-6.00\ g} = \frac{\Delta H\ (\text{butane})}{58.12\ g/mol}$$

ΔH (butane) = - 2890 kJ/mol (or 2890 kJ/mol released)

6. **B** These include the variables in the formula: $Q = mc\Delta T$ or $Q = mc(T_f - T_i)$

7. **A**

$$\begin{aligned}
-(Q_{loss}) &= Q_{gain} \\
-mc\,\Delta T &= mc\Delta T \quad \text{(but m and c are equal on both sides)} \\
-(T_f - T_i) &= (T_f - T_i) \\
-(T_f - 90.0^oC) &= (T_f - 25.0^oC) \\
-T_f + 90.0 &= T_f - 25.0 \\
-2\,T_f &= -25.0 - 90.0 \\
-2\,T_f &= -115.0 \\
T_f &= 57.5\ ^oC
\end{aligned}$$

8. $Q = mc\Delta T$

$Q = (24.6\text{ g})(0.39\text{ J/g}^{o}\text{C})(25\,^{o}\text{C} - 0\,^{o}\text{C})$

$Q = 2.4 \times 10^{2}\text{ J}$

9.

water	copper
$m = 230.0\text{ g}$	$m = 200.0\text{ g}$
$T_i = 25\,^{o}\text{C}$	$T_i = 47\,^{o}\text{C}$
	$c = 0.39\text{ J/g}^{o}\text{C}$

copper		water
$-(Q_{loss})$	=	Q_{gain}
$-mc\Delta T$	=	$mc\Delta T$
$-mc(T_f - T_i)$	=	$mc(T_f - T_i)$
$-(200.0\text{ g})(0.39\text{ J/g}^{o}\text{C})(T_f - 47^{o}\text{C})$	=	$(230.0\text{ g})(4.19\text{ J/g}^{o}\text{C})(T_f - 25^{o})$
$-(T_f - 47)$	=	$12.4(T_f - 25)$
$-(T_f - 47)$	=	$12.4\,T_f - 310$
T_f	=	$27\,^{o}\text{C}$

10. $$\frac{Q\text{ (water)}}{\Delta m\text{ (coal)}} = \frac{-\ \Delta H\text{ (coal)}}{\text{molar mass (coal)}}$$

$$\frac{Q\text{ (water)}}{0.600\text{ g}} = \frac{+\,391\text{ kJ/mol}}{12.01\text{ g/mol}}$$

$Q\text{ (water)} = +\,19.5\text{ kJ or } +\,19500\text{ J}$

$Q = mc\Delta T$

$+\,19500\text{ J} = (200.00\text{ g})(4.19\text{ J/g}^{o}\text{C})(T_f - 22.2\,^{o}\text{C})$

$T_f - 22.2 = 23.3$

$T_f = 45.5\,^{o}\text{C}$

11. a)

$$\begin{aligned}
-(Q_{loss}) &= Q_{gain} \\
-mc\Delta T &= mc\Delta T \\
-mc(T_f - T_i) &= mc(T_f - T_i) \\
-m(T_f - T_i) &= m(T_f - T_i) \\
-(350.0\text{ g})(T_f - 60.0^oC) &= (275.0\text{ g})(T_f - 22.0^oC) \\
-(T_f - 60.0) &= 0.786\,(T_f - 22.0) \\
-(T_f - 60.0) &= 0.786\,T_f - 17.29 \\
-1.786\,T_f &= -77.3 \\
T_f &= 43.3\ ^oC
\end{aligned}$$

b)

$$\begin{aligned}
-(Q_{loss}) &= Q_{gain} \\
-mc\Delta T &= mc\Delta T \\
-mc(T_f - T_i) &= mc(T_f - T_i) \\
-m(T_f - T_i) &= m(T_f - T_i) \\
-(300.0\text{ g})(T_f - 70.0^oC) &= (625\text{ g})(T_f - 43.3^oC) \\
-(T_f - 70.0) &= 2.08\,(T_f - 43.3) \\
-(T_f - 70.0) &= 2.08\,T_f - 90.1 \\
-3.08\,T_f &= -160.1 \\
\text{Final}\quad T_f &= 52.0\ ^oC
\end{aligned}$$

12.

$$\begin{aligned}
-(Q_{loss}) &= Q_{gain} \\
-mc\Delta T &= mc\Delta T \\
-mc(T_f - T_i) &= mc(T_f - T_i) \qquad \text{Assume } c_{coffee} = 4.19\ \text{J/g}^oC \\
-m(T_f - T_i) &= m(T_f - T_i) \\
-(125.0\text{ g})(T_f - 98^oC) &= (300.00\text{g})(T_f - 22^oC) \\
-(T_f - 98) &= 2.4\,(T_f - 22) \\
-(T_f - 98) &= 2.4\,T_f - 52.8 \\
-3.4\,T_f &= -150.8 \\
T_f &= 44\ ^oC
\end{aligned}$$

OBJECTIVE 2.5

1. **D** When equation 2) is reversed, and the net ΔH is found, the net equation is: $1/2\ Cl_2(l) \rightarrow 1/2\ Cl_2(g)$

2. **A** (the ΔH for the net reaction is negative, but since the question asks for the "heat released", the answer is given as a positive value.)

3. **A** The second equation is multiplied by 2, and the third equation is reversed, to give the net equation and net ΔH value.

4. 2 x, and reverse a): $2NO_{(g)} \rightarrow O_{2(g)} + N_{2(g)}$ $\Delta H = -180.6$ kJ

2 x b):	$N_{2(g)} + 2O_{2(g)} \rightarrow 2NO_{2(g)}$	$\Delta H = 67.6$ kJ
NET:	$2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$	$\Delta H = -113$ kJ

5.

5 x a):	$10PCl_{3(l)} + 5\,O_{2(g)} \rightarrow 10POCl_{3(l)}$	$\Delta H = -2935$ kJ
reverse b):	$10POCl_{3(l)} \rightarrow P_4O_{10(s)} + 6PCl_{5(s)}$	$\Delta H = +419$ kJ
5 x c):	$10P_{(s)} + 15Cl_{2(g)} \rightarrow 10PCl_{3(l)}$	$\Delta H = -3430$ kJ
3 x, and reverse d):	$6PCl_{5(s)} \rightarrow 6P_{(s)} + 15Cl_{2(g)}$	$\Delta H = +2676$ kJ
NET:	$4P_{(s)} + 5O_{2(g)} \rightarrow P_4O_{10(s)}$	$\Delta H = -3270$ kJ

6.

	$C_6H_4(OH)_{2(aq)} \rightarrow C_6H_4O_{2(aq)} + H_{2(g)}$	$\Delta H = +177.4$ kJ
	$H_2O_{2(aq)} \rightarrow H_{2(g)} + O_{2(g)}$	$\Delta H = +191.2$ kJ
	$2\,H_{2(g)} + O_{2(g)} \rightarrow 2\,H_2O_{(g)}$	$\Delta H = -485.6$ kJ
	$2\,H_2O_{(g)} \rightarrow 2\,H_2O_{(l)}$	$\Delta H = -87.6$ Kj
NET:	$C_6H_4(OH)_{2(aq)} + H_2O_{2(aq)} \rightarrow C_6H_4O_{2(aq)} + 2\,H_2O_{(l)}$	$\Delta H = -204.6$ kJ

MODULE 4
ANSWER KEY

OBJECTIVE 1.1

1. B — Gunpowder reacts in less than a second after impact.
Sodium in water will take several seconds depending on the size of the piece of sodium.
A compost pile may take several years to decompose.

OBJECTIVE 1.2

1. B — A rate is a change in a quantity per unit of time.

2. D — In one minute, 24 litres of NH_3 are formed. Since 3 H_2 are consumed to make 2 NH_3, 36 litres of H_2 per minute would be required to produce 24 litres of NH_3 per minute.

3. B — Number of moles of O_2 needed = 24 g ÷ 32 g/mole = 0.75 moles
Since 2 H_20 are required to produce 1 O_2, 1.5 moles of H_20 are required to produce 0.75 moles of O_2.
1.5 mole of H_2O x 18 g/mole = 27 g of H_2O needed
27 g of H_2O ÷ 1.0 g/mL = 27 mL of H_2O needed
27 mL ÷ 45.0 mL/h = 0.60 h
0.60 hours x 60 minutes/hour = 36 minutes

4. Rate of formation of Y_2 = 0.4 moles per litre per minute
From the graph, at 8 minutes, the concentration of X is 6.4 mol/L.
At 0 minutes, the concentration was 0 moles per litre per minute.
Therefore, in the first 8 minutes, the concentration of X increased by 6.4 moles per litre.
Since 1 mole of Y_2 is produced for every 2 moles of X (from the balanced chemical equation), an increase of 6.4 moles per litre of X would mean an increase of 3.2 moles per litre of Y_2.
Therefore, the average rate of formation of Y_2 is
3.2 moles per litre ÷ 8 minutes = 0.4 moles per litre per minute.

OBJECTIVE. 2.1

1. C Spontaneous combustion can result if a fuel is heated to a sufficiently high temperature to initiate the reaction between the fuel and oxygen.
   A source of spark or flame is not necessary.

2. D This choice produces the largest surface area and therefore the greatest rate of reaction.

3. B Although the foam contains some water which acts to cool the fire, the main purpose of the foam is to prevent oxygen from reaching the fuel.

OBJECTIVE 2.2

1. A A change in pressure only affects the rate of a reaction if gases are involved.

2. D Increasing the concentration increases the rate. Increasing the temperature also increases the rate. Therefore the greatest rate will occur at the highest temperature and largest concentration.

3. B Since the reactants in a precipitation reaction are in the aqueous state, the ions need only combine. Since no bonds need to be broken, the reaction is very fast. Although CH_4 burns rapidly in oxygen once ignited, it does not react rapidly at room temperature.

OBJECTIVE 2.3

1. B Since the curve was shifted to the left, the molecules had less average kinetic energy and therefore the temperature must have been lowered. Since the activation energy line also shifted to the left, a catalyst must have been added to lower the activation energy barrier.

2. C The reactants started at 40 kJ and the peak occurred at 90 kJ.
   Therefore, the activation energy is 90 kJ - 40 kJ = 50 kJ
   For the reverse reaction, the reactants would start at 60 kJ and the products would have an energy of 40 kJ.
   Therefore, the $\Delta H = 40 \text{ kJ} - 60 \text{ kJ} = -20 \text{ kJ}$

3. D A catalyst lowers the activation energy of both the forward and the reverse reactions, increasing the rate of both processes.

4. C A catalyst lowers the activation energy needed and therefore increases the rate by making collisions more likely to be effective, not by increasing the number of collisions between the molecules.

OBJECTIVES. 3.1 and 3.3

1. C Genetic alteration does not directly rely on modifying reaction rates. A larger tomato did not necessarily grow faster.

MODULE 5

ANSWER KEY

OBJECTIVES 1.1, 1.2, 2.1 and 2.2

1. B — all substances must be present at equilibrium

2. D — activity must be occurring on the microscopic level with no macroscopic activity

3. C — this is the only reversible process - some dissolved solute is precipitating while some undissolved solute (solid at the bottom) is dissolving

4. C — amounts of liquid benzene and benzene vapor are both constant - the system is closed and the rates of condensation and vaporization (evaporation) are equal

5. B — a catalyst has no effect once equilibrium has been attained - surface area affects only the rate of reaction - both of these factors influence initial rates but not equilibrium

6. A — the higher the K_C value, the more products are favored at equilibrium; the lower the K_C value, the less products are favored at equilibrium.

7. C — at times 5 and 6 minutes, the graph lines indicate constant concentration - of the times given, 5 minutes is the earliest

8. A — decreasing the temperature is the same as removing energy - the system will shift to the right - products will increase and reactants will decrease

9. A — The equilibrium concentrations of the products, raised to the appropriate powers, appear in the numerator of the K_C expression.

10. C — The reason is similar to #9 with one additional fact. A Kc expression does not include any concentrations which are fixed. Since the concentration of water, H2O$_{(l)}$ is fixed, it is not included.

11. D — same explanation as #10

12. C — $H_2O_{(l)}$ and solids are excluded for K_C

13. B — the K_C value will increase but one cannot find the actual value without more information

14. A $K_c = [N_2O_4]/[NO_2]^2 = (4.0)/(3.0)^2 = 0.44$

15. B $N_2O_4 \leftrightarrow 2\ NO_2$

I	2	0
C	-0.4	+0.8
E	1.6	0.8

$K_c = [NO_2]^2/[N_2O_4] = (0.8)^2/1.6 = 0.4$

16. B $2NO_2 \leftrightarrow 2NO + O_2$

I	5	0	0
C	-3	+3	+1.5
E	2	3	1.5

$K_c = [NO]^2[O_2]/[NO_2]^2 = (3)^2(1.5)/(2)^2 = 3.38$

17. D $A + 2\ B \leftrightarrow 2X + 4Y$

I	4	4	0	0
C	-1	-2	+2	+4
E	3	2	2	4

$K_c = [X]^2[Y]^4/[A][B]^2 = (2)^2(4)^4/(3)(2)^2 = 85.33$

18. C $N_2 + 2\ O_2 \leftrightarrow 2\ NO_2$

I	0.6	1.2	0
C	-0.1	-0.2	+0.2
E	0.5	1.0	0.2

$K_c = [NO_2]^2/[N_2][O_2]^2 = (0.2)^2/(0.5)(1.0)^2 = 0.08$

19. D K_c = [products]/[reactants], products of the reaction are CO_2 and H_2O and the reactants are C_3H_8 and O_2 with the exponents in K_c as coefficients in the equation

20. $K_c = [SO_3]^2/[SO_2]^2[O2] = (0.15)^2/(0.3)^2(0.05) = 5$

OBJECTIVE 1.3

1. B raising the pressure has no effect (4 moles of gas on each side of the equation) a catalyst has no effect once equilibrium is attained

2. A adding H^+ ions decreases OH^- (neutralization) - the equilibrium will shift to try to replace OH^- ions, therefore the solution becomes more orange

3. C 1- change will cause a shift to the right
2- change will have no effect
3- change will cause a shift to the left
4- pressure decreases; change will cause a shift to the right

4. A 1- favors the production of fewest gas particles; i.e., more NH_3
5- increasing reactant concentration(s) will increase product concentration(s); the equilibrium will shift to the right

5. B increase [CO_2] causes a decrease in energy, increase in CO and O_2

6. A - increase in pressure favors the side with fewer gas particles (5 reactant vs 2 product)
- removal of energy shifts the equilibrium to the right

7. A A) causes a shift to the right where B),C),and D) cause shifts to the left

8. B an increase in pressure favors the side with fewer gas particles; therefore PCl_5 concentration will increase

9. A adding $Cu(NO_3)_{2(s)}$ is the same as adding Cu^{2+} ions; the effect will be to shift the equilibrium to the right

10. D A) shifts to the left
B) no effect
C) shifts to the left
D) shifts to the right

11. a, c and f
    - A) favors the forward reaction (3 moles of reactant vs 2 moles of product)
    - B) no effect
    - C) less N_2, more NO_2 at the new equilibrium
    - D) removing energy; less NO_2, more N_2 and O_2
    - E) all concentrations will increase (equilibrium shifts to the left, but not all the extra NO_2 will react)
    - F) opposite effect of D)

12. Addition of $NaOH_{(s)}$ means the addition of OH^- ions. These ions will neutralize H^+ ions causing a decrease in $[H^+]$. This will cause a shift to the left. $[Pb^{2+}]$ and $[H_2]$ will decrease and more $Pb_{(s)}$ will appear.

13. Addition of Ag^+ decreases the $[SCN^-]$. This causes a shift to the left. $[Fe^{3+}]$ will increase, $[SCN^-]$ will decrease and $[FeSCN^{2+}]$ will decrease. The red of the $FeSCN^{2+}$ ion will fade and the yellow of the Fe^{3+} ion will intensify.

14. Increasing the temperature will cause more AB to form and less A_2 and B_2 to be present at the new equilibrium. Diagram A is the ONLY one showing an increase in AB and decreases in A_2 and B_2.

15. Increasing the temperature will favor the decomposition of HI (shift to the left). $[H_2]$ and $[I_2]$ will increase and [HI] will decrease. Since $\boldsymbol{K_c}$ = [products]/[reactants], its value will decrease.

OBJECTIVE 2.3

1. $[H^+_{(aq)}] = 1 \times 10^{-7}$ mol/L
   $[OH^-_{(aq)}] = 1 \times 10^{-7}$ mol/L
2. $K_W = 1 \times 10^{-14}$
3. B pH is a logarithmic scale in which 100 is represented by a pH of 2, that is 2 less than water with pH = 7
4. B Acids dissociate to release H^+ ions in solution.

OBJECTIVE 2.5

1. a) 1.0×10^{-11} mol/L
   b) 1.0×10^{-8} mol/L
   c) 1.0×10^{-2} mol/L

2. $[H^+][OH^-] = 1.0 \times 10^{-14}$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-3}} = 1.0 \times 10^{-11} \text{ mol/L}$$

3. The greater the K_a the greater the degree of ionization, and the stronger the acid. Increasing order from top to bottom.

HCO_3^- (aq)	$\leftrightarrow$	$H^+(aq) + CO_3^{2-}(aq)$	$K_a = 5.6 \times 10^{-11}$
$NH_4^+(aq)$	$\leftrightarrow$	$H^+(aq) + NH_3(aq)$	$K_a = 5.6 \times 10^{-10}$
$HF(aq)$	$\leftrightarrow$	$H^+(aq) + F^-(aq)$	$K_a = 3.5 \times 10^{-4}$
$HNO_2(aq)$	$\leftrightarrow$	$H^+(aq) + NO_2^-(aq)$	$K_a = 4.6 \times 10^{-4}$

4. The greater the K_a the greater the degree of ionization, and the stronger the acid Increasing order from top to bottom.

$HClO_{(aq)}$	$K_a = 1.0 \times 10^{-8}$
$HClCH_2CO_{2(aq)}$	$K_a = 1.4 \times 10^{-3}$
$HClO_{3(aq)}$	$K_a = 1.2 \times 10^{-2}$

5. $HF_{(aq)} \leftrightarrow H^+_{(aq)} + F^-_{(aq)}$

$pH = - \log [H^+]$
$5.50 = - \log [H^+]$
$[H^+] = 3.16 \times 10^{-6}$ mol/L

	$HF_{(aq)}$ $\leftrightarrow$	$H^+_{(aq)}$ +	$F^-_{(aq)}$
Start	0.100 mol/L	0	0
Shift	-x	+ x	+ x
Equilibrium	0.100 - x $0.100 - 3.16 \times 10^{-6}$ 0.0999 mol/L	3.16×10^{-6} mol/L	3.16×10^{-6} mol/L

$$K_a = \frac{[H^+_{(aq)}][F^-_{(aq)}]}{[HF_{(aq)}]} = \frac{(3.16 \times 10^{-6})(3.16 \times 10^{-6})}{(0.0999)} = 1.0 \times 10^{-10}$$

6. D

	$HF_{(aq)}$ $\leftrightarrow$	$H^+_{(aq)}$ +	$F^-_{(aq)}$
Start	2.0 mol/L	0	0
Shift	-x	+ x	+ x
Equilibrium	2.0 - x	x	x

$$K_a = \frac{[H^+_{(aq)}][F^-_{(aq)}]}{[HF_{(aq)}]} = \frac{(x)(x)}{2.0 - x} = 6.7 \times 10^{-4}$$

$x^2 = 13.4 \times 10^{-4} - (6.7 \times 10^{-4} x)$
$x^2 + (6.7 \times 10^{-4} x) - 13.4 \times 10^{-4} = 0$
Using any method to solve a quadratic equation
$x = 0.0363 = [H^+_{(aq)}] = 3.6 \times 10^{-2}$ mol/L

7.

$$\text{Concentration of } [OH^-_{(aq)}] = \frac{0.4 \text{ g}}{40 \text{ g/mol}} / 1 \text{ L}$$

$[OH^-_{(aq)}] = 0.01$ mol/L

$[H^+] \times [0H^-] = 1.0 \times 10^{-14}$

$$[H^+] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-2}} = 1.0 \times 10^{-12} \text{ mol/L}$$

pH = 12

8. $HF_{(aq)} \leftrightarrow H^+_{(aq)} + F^-_{(aq)}$

	$HF_{(aq)}$ $\leftrightarrow$	$H^+_{(aq)}$ +	$F^-_{(aq)}$
Start	0.25 mol/L	0	0
Shift	-x	+ x	+ x
Equilibrium	0.25 - x $0.25 - 4.0 \times 10^{-3}$ 0.246 mol/L	4.0×10^{-3} mol/L	4.0×10^{-3} mol/L

$$K_a = \frac{[H^+_{(aq)}][F^-_{(aq)}]}{[HF_{(aq)}]} = \frac{(4.0 \times 10^{-3})(4.0 \times 10^{-3})}{(0.246)} = 6.5 \times 10^{-5}$$

9. $NH4OH_{(aq)} \leftrightarrow NH4^+_{(aq)} + OH^-_{(aq)}$

$pH = -\log[H^+]$
$10 = -\log[H^+]$
$[H^+] = 1.0 \times 10^{-10}$ mol/L
$[OH^-] = 1.0 \times 10^{-4}$ mol/L

	$NH_4OH_{(aq)}$ $\leftrightarrow$	$NH_4^+_{(aq)}$ +	$OH^-_{(aq)}$
Start	0.12 mol/L	0	0
Shift	-x	+ x	+ x
Equilibrium	0.12 - x $0.12 - 1.0 \times 10^{-4}$ 0.1199 mol/L	1.0×10^{-4} mol/L	1.0×10^{-4} mol/L

$$K_a = \frac{[NH_4^+_{(aq)}][OH^-_{(aq)}]}{[NH_4OH_{(aq)}]} = \frac{(1.0 \times 10^{-4})(1.0 \times 10^{-4})}{(0.1199)} = 8.3 \times 10^{-8}$$

10. $pH = -\log[H^+]$
$3.50 = -\log[H^+]$
$[H^+] = 3.16 \times 10^{-4}$ mol/L

$$[OH^-] = \frac{1.0 \times 10^{-14}}{3.16 \times 10^{-4}} = 3.16 \times 10^{-11} \text{ mol/L}$$

11.

Acid	Equilibrium conc. of acid (mol/L)	pH	K_a
HOCl	0.30	4	**3.3×10^{-8}**
$HC0_2H$	**5.6×10^{-1}**	2	1.8×10^{-4}
HOBr	5.0×10^{-2}	**5**	2.0×10^{-9}

$HOCl_{(aq)} \leftrightarrow H^+_{(aq)} + OCl^-_{(aq)}$

pH = - log [H^+]
4.0 = - log [H^+]
[H^+] = 1.0×10^{-4} mol/L

$$K_a = \frac{[H^+_{(aq)}]\,[OCl^-_{(aq)}]}{[HOCl_{(aq)}]} = \frac{(1.0 \times 10^{-4})\,(1.0 \times 10^{-4})}{0.3} = 3.3 \times 10^{-8}$$

$HCO_2H_{(aq)} \leftrightarrow H^+_{(aq)} + CO_2H^-_{(aq)}$

pH = 2
[H^+] = 1.0×10^{-2} mol/L

$$K_a = \frac{[H^+_{(aq)}]\,[C0_2H^-_{(aq)}]}{[HC0_2H_{(aq)}]} = \frac{(1.0 \times 10^{-2})\,(1.0 \times 10^{-2})}{x} = 1.8 \times 10^{-4}$$

$x = 0.56 = [H^+_{(aq)}] = 5.6 \times 10^{-1}$ mol/L

$HOBr_{(aq)} \leftrightarrow H^+_{(aq)} + OBr^-_{(aq)}$

$$K_a = \frac{[H^+_{(aq)}]\,[OBr^-_{(aq)}]}{[HOBr_{(aq)}]} = \frac{(x)\,(x)}{(5.0 \times 10^{-2})} = 2.0 \times 10^{-9}$$

$x^2 = (5.0 \times 10^{-2})(2.0 \times 10^{-9}) = 1.0 \times 10^{-10}$

$x = \sqrt{1.0 \times 10^{-10}} = 1.0 \times 10^{-5}$ mol/L

$pH = -\log(1.0 \times 10^{-5}) = 5$

OBJECTIVE 3.1

1. $Al_{(s)} + Fe^{3+}{}_{(aq)} \rightarrow Fe^{2+}{}_{(aq)} + Al^{3+}{}_{(aq)}$

   Reduction equation: $Fe^{3+}{}_{(aq)} + 1e^- \rightarrow Fe^{2+}{}_{(aq)}$ (gains electron)
   Oxidation equation: $Al_{(s)} \rightarrow Al^{3+}{}_{(aq)} + 3e^-$ (loses electrons)
   Reducing agent: Al(s) (loses electrons)
   Oxidizing agent: $Fe^{3+}{}_{(aq)}$ (gains electron)

2. $Mg_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Mg^{2+}{}_{(aq)} + Cu_{(s)}$

   Reduction equation: $Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ (gains electrons)
   Oxidation equation: $Mg_{(s)} \rightarrow Mg^{2+}{}_{(aq)} + 2e^-$ (loses electrons)
   Reducing agent: $Mg_{(s)}$ (loses electrons)
   Oxidizing agent: $Cu^{2+}{}_{(aq)}$ (gains electrons)

3. $Au^{3+}{}_{(aq)} + Cd_{(s)} \rightarrow Au_{(s)} + Cd^{2+}{}_{(aq)}$

   Reduction equation: $Au^{3+}{}_{(aq)} + 3e^- \rightarrow Au_{(s)}$ (gains electrons)
   Oxidation equation: $Cd_{(s)} \rightarrow Cd^{2+}{}_{(aq)} + 2e^-$ (loses electrons)
   Reducing agent: Cd(s) (loses electrons)
   Oxidizing agent: $Au^{3+}{}_{(aq)}$ (gains electrons)

4. D

5. D
   Zinc loses $2e^-$ and undergoes oxidation, Zn is the reducing agent.
   $Cu^{2+}{}_{(aq)}$ gains $2e^-$ and undergoes reduction, $Cu^{2+}{}_{(aq)}$ is the oxidizing agent.

6. D

7. A

8. Since the $Al(NO_3)_{3(aq)}$ reacted with none of the metals, Al^{3+} ions have the smallest tendency to accept electrons and undergo reduction. $Zn(NO_3)_{2(aq)}$ reacted with only one metal, meaning that Zn^{2+} ions have the second smallest tendency to undergo reduction. $Ni(NO_3)_{2(aq)}$ reacted with two different metals, placing Ni^{2+} ions third. $Cu(NO_3)_{2(aq)}$ reacted with Ni, Zn and Al and therefore Cu2+ ions had he greatest tendency to undergo reduction.

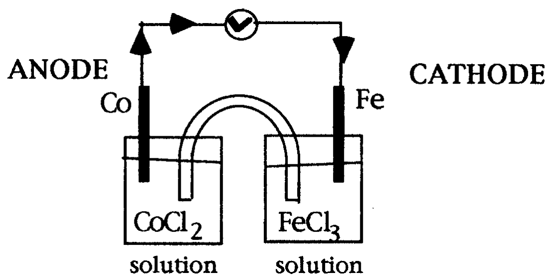
9. Oxidation half-reaction: half-reaction in which electrons are released
$Cr_{(s)} \rightarrow Cr^{3+}_{(aq)} + 3e^-$ and $2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$

10. C

11. A

OBJECTIVE 3.2

1.



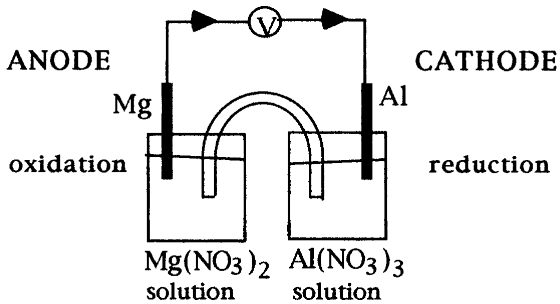
Since the arrow points to the right, electrons are generated at the Co electrode (ANODE) and travel towards the Fe electrode (CATHODE).
The balanced half reactions for this cell are:

$$3(Co_{(s)} \rightarrow Co^{2+}{}_{(aq)} + 2e^-) \qquad \text{oxidation}$$

$$2(Fe^{3+}{}_{(aq)} + 3e^- \rightarrow Fe_{(s)}) \qquad \text{reduction}$$

NET: $2Fe^{3+}{}_{(aq)} + 3Co_{(s)} \rightarrow 3Co^{2+}{}_{(aq)} + 2Fe_{(s)}$

2.



Since the electrons are lost by the Mg electrode (ANODE), the Al electrode (CATHODE) is accepting electrons. The cathode is where reduction occurs.

The half-reaction at the cathode is: $2(Al^{3+}{}_{(aq)} + 3e^- \rightarrow Al_{(s)})$

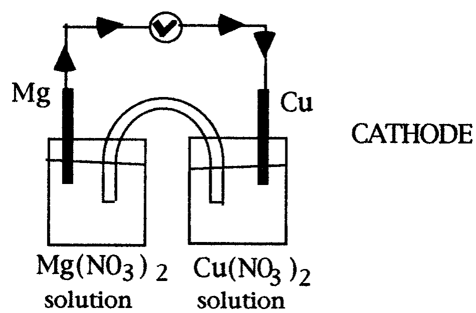
The half-reaction at the anode is: $3(Mg_{(s)} \rightarrow Mg^{2+}{}_{(aq)} + 2e^-)$

The complete balanced equation for the net equation is:

$$2Al^{3+}{}_{(aq)} + 3Mg_{(s)} \rightarrow 3Mg^{2+}{}_{(aq)} + 2Al_{(s)}$$

The electrode which increases in mass is the Al electrode because Al^{3+} ions in the solution are accepting electrons and are plating on the Al electrode.

3. Cathode is where reduction takes place. This is the electrode which accepts electrons, therefore the electrons are originating at the Mg electrode and the current flows is from left to right.



4. B

$Sn_{(s)} \rightarrow Sn^{2+}{}_{(aq)} + 2e^-$ oxidation

$Ag^+{}_{(aq)} + 1e^- \rightarrow Ag_{(s)}$ reduction

Oxidation (loss of electrons) takes place at the anode which is the Sn electrode. This is the electrode which loses mass.
Reduction (gain of electrons) takes place at the cathode which is the Ag electrode. This is the electrode which gains mass.

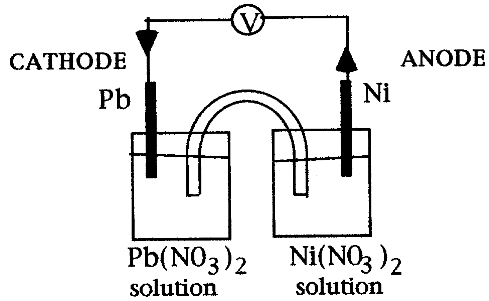
OBJECTIVE 3.3

1. Using the standard reduction potential table;

$Pb^{2+}_{(aq)} + 2e^- \rightarrow Pb_{(s)}$ $\quad E° = -0.13$ V $\quad$ reduction

$Ni^{2+}_{(aq)} + 2e^- \rightarrow Ni_{(s)}$ $\quad E° = -0.26$ V

Reduction takes place at the CATHODE and oxidation at the ANODE.



2. Using the standard reduction potential table;

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ $\quad E° = +0.34$ V $\quad$ reduction

$Al^{3+}_{(aq)} + 3e^- \rightarrow Al_{(s)}$ $\quad E° = -1.66$ V

$3(Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)})$ $\quad E° = +0.34$ V $\quad$ reduction

$2(Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^-)$ $\quad E° = +1.66$ V $\quad$ oxidation

NET: $3Cu^{2+}_{(aq)} + 2Al_{(s)} \rightarrow 2Al^{3+}_{(aq)} + 3Cu_{(s)}$ $\quad E° = +2.00$ V

3. Using the standard reduction potential table;

$Au^{3+}_{(aq)} + 3e^- \rightarrow Au_{(s)}$ $\quad E° = +1.50$ V

$Al^{3+}_{(aq)} + 3e^- \rightarrow Al_{(s)}$ $\quad E° = -1.66$ V

$Au^{3+}_{(aq)} + 3e^- \rightarrow Au_{(s)}$ $\quad E° = +1.50$ V $\quad$ reduction

$Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^-$ $\quad E° = +1.66$ V $\quad$ oxidation

NET: $Au^{3+}_{(aq)} + Al_{(s)} \rightarrow Al^{3+}_{(aq)} + Au_{(s)}$ $\quad E° = +3.16$ V

4. $Cr_{(s)} \mid Cr^{2+}{}_{(aq)} \parallel Cu^{2+}{}_{aq)} \mid Cu_{(s)}$

Using the standard reduction potential table;

$Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ $E° = +0.34\ V$
$Cr^{2+}{}_{(aq)} + 2e^- \rightarrow Cr_{(s)}$ $E° = -0.91\ V$

$Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ $E° = +0.34\ V$ reduction
$Cr_{(s)} \rightarrow Cr^{2+}{}_{(aq)} + 2e^-$ $E° = +0.91\ V$ oxidation

NET: $Cu^{2+}{}_{(aq)} + Cr_{(s)} \rightarrow Cr^{2+}{}_{(aq)} + Cu_{(s)}$ $E° = +1.25\ V$

The electrode which increases in mass: copper
The electrode which decreases in mass: chromium
The cell voltage: 1.25 V

5. Using the standard reduction potential table;

a) $Co^{2+}_{(aq)} + Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + Co_{(s)}$

$Co^{2+}_{(aq)} + 2e^- \rightarrow Co_{(s)}$ E° = -0.28 V reduction
$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$ E° = +0.44 V oxidation

NET: $Co^{2+}_{(aq)} + Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + Co_{(s)}$ E° = +0.16 V <u>spontaneous</u>

b) $Cu_{(s)} + 2H^+_{(aq)} \rightarrow Cu^{2+}_{(aq)} + H_{2(g)}$

$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$ E° = -0.34 V oxidation
$2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$ E° = 0.00 V reduction

NET: $Cu_{(s)} + 2H^+_{(aq)} \rightarrow Cu^{2+}_{(aq)} + H_{2(g)}$ E° = -0.34 V <u>not spontaneous</u>

c) $2Ag_{(s)} + Fe^{2+}_{(aq)} \rightarrow 2Ag^+_{(aq)} + Fe_{(s)}$

$2(Ag_{(s)} \rightarrow Ag^+_{(aq)} + 1e^-)$ E° = -0.80 V oxidation
$Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$ E° = -0.44 V reduction

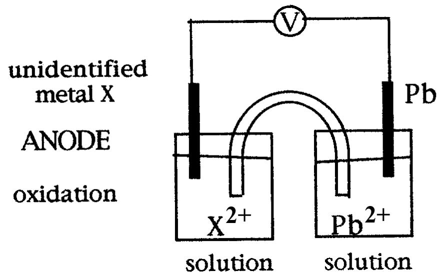
NET: $2Ag_{(s)} + Fe^{2+}_{(aq)} \rightarrow 2Ag^+_{(aq)} + Fe_{(s)}$ E° = -1.24 V <u>not spontaneous</u>

d) $3Zn^{2+}_{(aq)} + 2Cr_{(s)} \rightarrow 3Zn_{(s)} + 2Cr^{3+}_{(aq)}$

$3(Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)})$ E° = -0.76 V reduction
$2(Cr_{(s)} \rightarrow Cr^{3+}_{(aq)} + 3e^-)$ E° = +0.74 V oxidation

NET: $3Zn^{2+}_{(aq)} + 2Cr_{(s)} \rightarrow 3Zn_{(s)} + 2Cr^{3+}_{(aq)}$ E° = -0.02 V <u>not spontaneous</u>

6.



$Pb^{2+}{}_{(aq)} + 2e^- \rightarrow Pb_{(s)}$ $E° = -0.13$ V reduction

$X_{(s)} \rightarrow X^{2+}{}_{(aq)} + 2e^-$ $E° = ?$ oxidation

NET: $Pb^{2+}{}_{(aq)} + X_{(s)} \rightarrow X^{2+}{}_{(aq)} + Pb_{(s)}$ $E° = +0.15$ V

Since –0.13V + $E°_X$ = 0.15V, the standard oxidation potential of the unknown metal has to be 0.28 V. Using the Standard Reduction Potential table, the unknown metal must be cobalt (Co).

OBJECTIVE 3.4

1. We need a combination of metals which gives the highest electric potential, and over 1.5 V. The combination with a voltage greater than 1.5 V is between silver and zinc (if Ag is reduced and Zn is oxidized).

$2(Ag^+{}_{(aq)} + 1e^- \rightarrow Ag_{(s)})$ $E° = +0.80$ V

$Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2e^-$ $E° = +0.76$ V

$2Ag^+{}_{(aq)} + Zn_{(s)} \rightarrow 2Ag_{(s)} + Zn^{2+}{}_{(aq)}$ $E° = 1.56$ V

2. If the scientist places the chromium solution in the copper container then an oxidation-reduction reaction may occur.
That is:

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ $\quad$ E° = +0.34 V

$Cr^{3+}_{(aq)} + 3e^- \rightarrow Cr_{(s)}$ $\quad$ E° = -0.74 V

$3(Cu^{3+}_{(aq)} + 3e^- \rightarrow Cu_{(s)})$ $\quad$ E° = +0.34 V reduction

$2(Cr_{(s)} \rightarrow Cr^{3+}_{(aq)} + 3e^-)$ $\quad$ E° = +0.74 V oxidation

$$3Cu^{2+}_{(aq)} + 2Cr_{(s)} \underset{\leftarrow \text{ not spontaneous}}{\overset{\text{spontaneous } \rightarrow}{\leftrightarrow}} 3Cu_{(s)} + 2Cr^{3+}_{(aq)} \qquad E° = 1.08 \text{ V}$$

The positive cell potential favours the forward reaction. As a result, the reverse reaction cannot occur spontaneously and copper will not oxidize.
This copper container <u>can</u> therefore be used to store the $Cr(NO_3)_{3(aq)}$ solution.

APPENDIX I

FORMULAS

$Q = mc\Delta T$

$PV = n RT$

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

PHYSICAL CONSTANTS

SYMBOL	NAME	VALUE
c_{H_2O}	Specific heat capacity of water	4190 J/(kg•ºC) or 4.19 J/(g•ºC)
ρ_{H_2O}	Density of water	1.00 g/mL
R	Molar gas constant	8.31 L•kPa/(mol•K)

APPENDIX II

STANDARD REDUCTION POTENTIALS

ION CONCENTRATION of 1 mol/L at 25°C and 101.3 kPa.

Reduction Half-reaction				Reduction Potential (V)
$F_2(g)$	$+ 2e^-$	$\rightarrow$	$2F^-(aq)$	E° = + 2.87
$Au^{3+}(aq)$	$+ 3e^-$	$\rightarrow$	$Au(s)$	E° = + 1.50
$Cl_2(g)$	$+ 2e^-$	$\rightarrow$	$2Cl^-(aq)$	E° = + 1.36
$Br_2(aq)$	$+ 2e^-$	$\rightarrow$	$2Br^-(aq)$	E° = + 1.09
$Br_2(l)$	$+ 2e^-$	$\rightarrow$	$2Br^-(aq)$	E° = + 1.07
$Ag^+(aq)$	$+ e^-$	$\rightarrow$	$Ag(s)$	E° = + 0.80
$Hg^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Hg(l)$	E° = + 0.78
$Fe^{3+}(aq)$	$+ e^-$	$\rightarrow$	$Fe^{2+}(aq)$	E° = + 0.77
$I_2(s)$	$+ 2e^-$	$\rightarrow$	$2I^-(aq)$	E° = + 0.53
$Cu^+(aq)$	$+ e^-$	$\rightarrow$	$Cu(s)$	E° = + 0.52
$Cu^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Cu(s)$	E° = + 0.34
$2H^+(aq)$	$+ 2e^-$	$\rightarrow$	$H_2(g)$	E° = + 0.00
$Pb^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Pb(s)$	E° = - 0.13
$Sn^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Sn(s)$	E° = - 0.14
$Ni^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Ni(s)$	E° = - 0.26
$Co^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Co(s)$	E° = - 0.28
$Fe^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Fe(s)$	E° = - 0.44
$Cr^{3+}(aq)$	$+ 3e^-$	$\rightarrow$	$Cr(s)$	E° = - 0.74
$Zn^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Zn(s)$	E° = - 0.76
$Cr^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Cr(s)$	E° = - 0.91
$Mn^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Mn(s)$	E° = - 1.18
$Al^{3+}(aq)$	$+ 3e^-$	$\rightarrow$	$Al(s)$	E° = - 1.66
$Be^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Be(s)$	E° = - 1.85
$Mg^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Mg(s)$	E° = - 2.37
$Na^+(aq)$	$+ e^-$	$\rightarrow$	$Na(s)$	E° = - 2.71
$Ca^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Ca(s)$	E° = - 2.87
$Sr^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Sr(s)$	E° = - 2.89
$Ba^{2+}(aq)$	$+ 2e^-$	$\rightarrow$	$Ba(s)$	E° = - 2.91
$Cs^+(aq)$	$+ e^-$	$\rightarrow$	$Cs(s)$	E° = - 2.92
$K^+(aq)$	$+ e^-$	$\rightarrow$	$K(s)$	E° = - 2.93
$Rb^+(aq)$	$+ e^-$	$\rightarrow$	$Rb(s)$	E° = - 2.98
$Li^+(aq)$	$+ e^-$	$\rightarrow$	$Li(s)$	E° = - 3.04

APPENDIX III

PERIODIC TABLE OF THE ELEMENTS

Legend

Element symbol — Atomic number — Atomic mass

1 **H** 1.01

TRANSITION ELEMENTS

	I A 1	II A 2	III B 3	IV B 4	V B 5	VI B 6	VII B 7	VIII B 8	9	10	I B 11	II B 12	III A 13	IV A 14	V A 15	VI A 16	VII A 17	VIII A 18
1	1 **H** **1.01**																	2 **He** 4.00
2	3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
3	11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
4	19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.71	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
5	37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98.91)	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.40	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.30
6	55 **Cs** 132.91	56 **Ba** 137.33	57-71 **La-Lu**	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.85	75 **Re** 186.20	76 **Os** 190.21	77 **Ir** 192.22	78 **Pt** 195.09	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.37	82 **Pb** 207.20	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
7	87 **Fr** (223)	88 **Ra** (226)	89-103 **Ac-Lr**	104 **Rf** (261)	105 **Ha** (262)	106 **Sg** (263)	107 **Uns** (262)	108 (265)	109 **Une** (266)	110 **Uun** (272)								

6	57 **La** 143.91	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.35	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.04	71 **Lu** 174.97
7	89 **Ac** 227.03	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** 237.05	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (254)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (260)

( ) The numbers in brackets represent the mass number of the most stable or most abundant isotope. The atomic masses of most of the elements are given to the nearest hundredth.

APPENDIX II