Review Book

Surviving Chemistry

One Concept at a Time

A ***Review*** of High School Chemistry Concepts

2012 Revision

e3Chemistry

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Surviving Chemistry One Concept at a Time

Review Book – 2012 Revision

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Table of Contents

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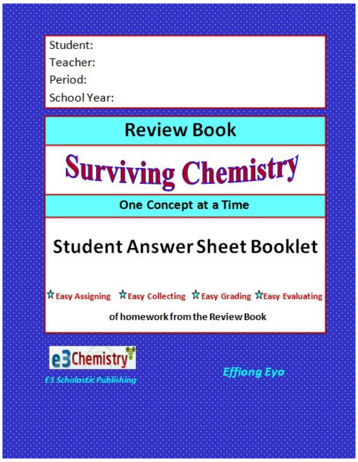

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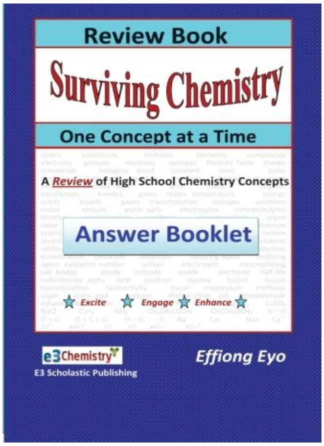

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A *Review* of High School Chemistry Concepts

Answer Booklet

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Lesson 1: Types of Matter

Introduction:

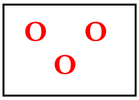
Chemistry is the study of matter; its composition, structures, properties, changes it undergoes, and energy accompanying these changes.

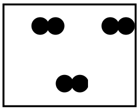
Matter is anything that has mass and takes up space. Matter, other words, is "stuff." Matter can be grouped and classified as pure substances or mixtures. In this lesson, you will learn about the different classifications of matter.

Types of Matter

Pure substances are types of matter composed (made up) of particles that are the same. Composition of a pure substance is uniform and definite in every sample. Elements and Compounds are classified as pure substances.

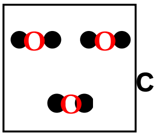
Elements are pure substances that are composed of identical atoms with the same atomic number. Elements cannot be decomposed (broken down) into simpler substances by physical nor chemical methods. Ca(s) and Br$_2$(g) are examples of elements. All known natural and synthesized elements can be found on the Periodic Table of the Elements.





elements Ca and Br

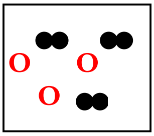
Compounds are pure substances composed of two or more different elements that are *chemically* combined. Properties and composition of a compound is definite (the same) in all samples of the compound. Compounds can be decomposed (separated) into simpler substances by chemical methods only. Properties of a compound are always different from those of the elements found in the compound. CaBr$_2$(s), H_2O(l), and NH_3(g) are examples of compounds.



a compound of
Ca and Br

Law of definite composition states that elements in a compound are combined in a fixed and definite ratio by mass. For example, the composition (mass percentages) in every sample of water is always 89% O to 11% H. That means any 10-gram sample of water will always contain about 8.9 g of O to 1.1g of H.

Mixtures are types of matter that are composed of two or more different substances that are *physically* combined. Composition of a mixture may vary (can change) from one sample to another. A mixture can be separated into its components only by physical methods. A mixture always retains the properties of the individual component.



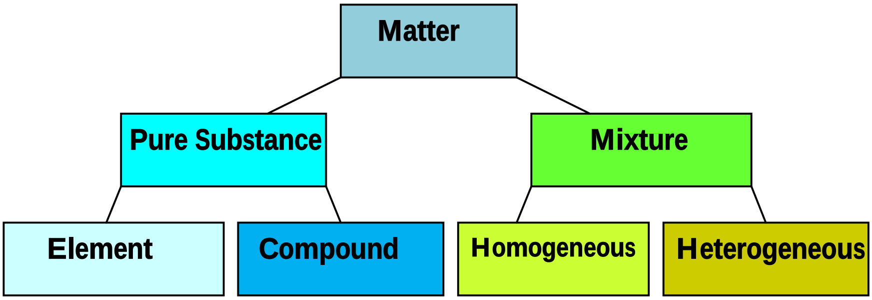
a mixture of
Ca and Br

Homogeneous and heterogeneous mixtures

Homogeneous mixtures are mixtures that are uniformly and evenly mixed throughout. Samples taken within the same mixture have definite and fixed composition. Aqueous solutions are homogeneous mixtures that are made with water. Salt water, NaCl(aq), is an example of aqueous.

Heterogeneous mixtures are mixtures that are not uniformly nor evenly mixed throughout. Samples taken within the same mixture have different and varying compositions. Soil and concrete are examples of heterogeneous.

Classification of matter diagram



Separation of mixtures

In a mixture, substances retain their unique physical properties. Depending on these physical properties, various physical methods can be used to separate each substance from the mixture.

Heterogeneous mixtures can be separated by simple physical methods.
Decantation is a process of pouring out the top component of a mixture that has separated into layers. An oil and water mixture can be separated this way.

Filtration is a process that can be used to separate a solid from liquid or aqueous. During filtration, the liquid or aqueous component of a mixture will go through the filter paper because particles of a liquid are always smaller than the holes of a filter. The solid component of the mixture will remain on the filter paper because particles of a solid are generally bigger than the holes of a filter.

Homogeneous mixtures (such as solutions) can be separated by more complicated physical methods.
Distillation is a process of separating a homogeneous mixture (solution) by using differences in the boiling points of the substances in the mixture. During distillation, a mixture is heated to vaporize (boil off) each substance in the order from lowest to highest boiling point. Each substance can be condensed and collected as it leaves the mixture. Water can be separated from salt in a salt-water mixture by simply boiling and evaporating the water off in a simple distillation apparatus. A mixture of hydrocarbons (methane, ethane, propane..etc) can be separated through a more complicated distillation process.
Chromatography is another method of separating homogeneous mixtures. In this process, a mixture is dissolved in a solvent (mobile phase) that allows the components of the mixture to move though a stationary phase at different speeds. Data from chromatograph separation can be collected and analyzed to learn about the mixture.

Lesson 2: Phases of Matter, Energy and Temperature

Introduction

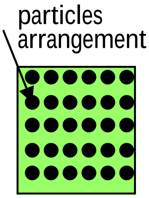
There are three phases of matter: solid, liquid, and gas. The nature of a substance determines the phase in which the substance will exist under normal conditions. Most substances can change from one phase to another. The nature of a substance also determines conditions necessary for the substance to change from one phase to another.

In this lesson you will learn about the three phases of matter. You will also learn about phase changes of matter and the relationship to temperature and energy.

Phases of Matter

Solid: A substance in the solid phase has the following characteristics:

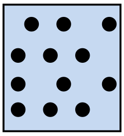
- Definite volume and definite shape
- Particles arranged orderly in a "*regular geometric pattern*"
- Particles vibrating around fixed points
- Particles with strong attractive forces to one another
- Particles that cannot be easily compressed (incompressible)



$H_2O(s)$

Liquid: A substance in the liquid phase has the following characteristics:

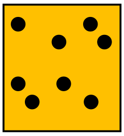
- Definite volume, but no definite shape (it takes the shape of its container)
- Particles that flow over each other
- Particles that cannot be easily compressed (incompressible)



$H_2O(l)$

Gas: A substance in the gas phase has the following characteristics:

- No definite volume and no definite shape (it takes volume and shape of its container)
- Particles that are less orderly arranged (most random)
- Particles that move fast and freely
- Particles with very weak attractive forces to each other
- Particles that can be easily compressed (compressible)



$H_2O(g)$

Phase changes

A **phase change** is a physical change. During a phase change, a substance changes its form (or state) without changing its chemical composition. Any substance can change from one phase to another given the right conditions of temperature and/or pressure. Most substances require a large change in temperature to go through one phase change. Water is one of a few chemical substances that can change through all three phases within a narrow range of temperature changes.

Phase changes and example equations representing each change are given below.

Melting is a change from *solid* to *liquid.* $H_2O(s)$ --- > $H_2O(l)$

Freezing is a change from *liquid* to *solid* $H_2O(l)$ ---> $H_2O(s)$

Evaporation is a change from *liquid* to *gas* $C_2H_5OH(l)$ ---> $C_2H_5OH(g)$

Condensation is a change from *gas* to *liquid* $C_2H_5OH(g)$ ---> $C_2HOH(l)$

Deposition is a change from *gas* to *solid* $CO_2(g)$ -----> $CO_2(s)$

Sublimation is a change from *solid* to *gas* $CO_2(s)$ ----> $CO_2(g)$

Iodine, $I_2(s)$ and dry ice, $CO_2(s)$, are two substances that readily sublime at normal conditions. Most substances do not sublime.

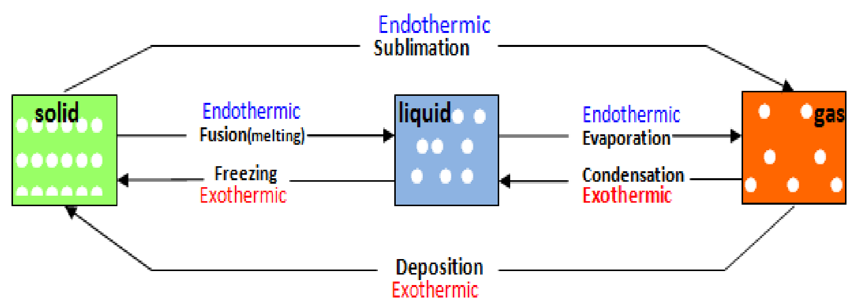
Phase changes and energy

A substance changes phase when it has absorbed or released enough heat energy to rearrange its particles (atoms, ions, or molecules) from one form to another. Some phase changes require a release of heat by the substance, while others require heat to be absorbed.

Endothermic describes a process that absorbs heat energy.
Fusion, evaporation and sublimation are endothermic phase changes.

Exothermic describes a process that releases heat energy.
Freezing, condensation and deposition are exothermic phase changes.

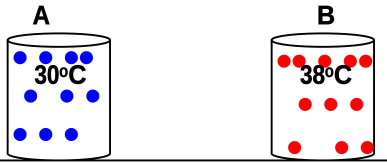
The diagram below summarizes phase changes and the relationship to energy.



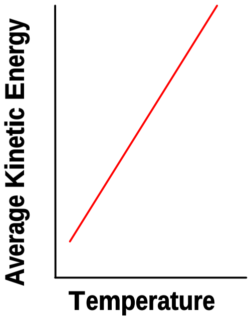
Phase changes and Temperature

A phase change for a substance occurs at a specific temperature. Every substance has its own unique melting and boiling point.

Temperature is a measure of the average kinetic energy of particles in matter.
Kinetic energy is energy due to movements of particles in matter . The higher the temperature of a substance, the greater its kinetic energy. As temperature increases, the average kinetic energy also increases.



Since particles ● in **B** are at a higher temperature, ● will be moving faster (higher kinetic energy) than particles ● in **A**



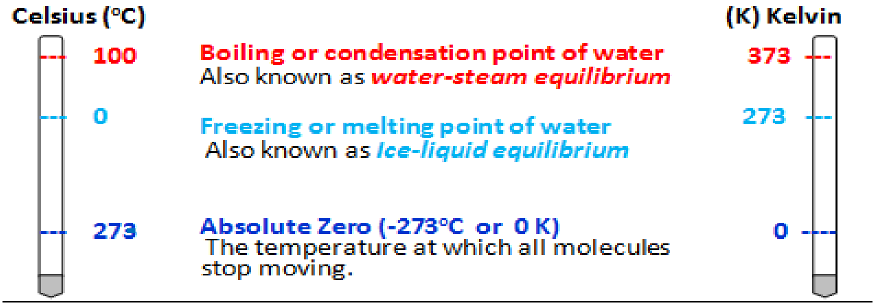
Thermometer is a piece an equipment that is used for measuring temperature. **Degree Celsius (°C) and Kelvin (K)** are the two most common units for measuring temperature.

Two fixed reference points are needed to create a thermometer scale:
The *freezing point (0°C , 273K)* and the *boiling point (100°C , 373K)* of water are often used as the two reference points in creating thermometer scales.
The mathematical relation between Celsius and Kelvin is given below.

$$K = {}^{o}C + 273$$ Table T equation

According to this equation, the Kelvin temperature value is always 273 units higher than the same temperature in Celsius.

Important temperature points



Boiling or condensation point of water
Also known as ***water-steam equilibrium***

Freezing or melting point of water
Also known as ***Ice-liquid equilibrium***

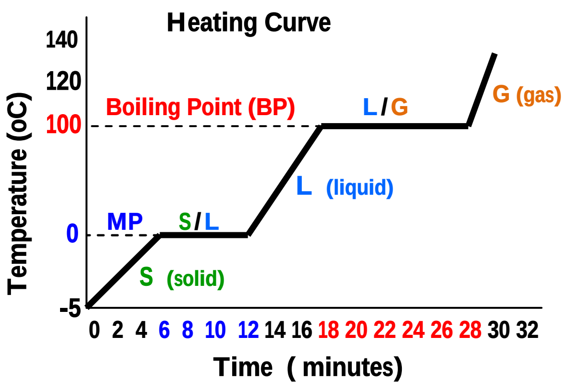
Absolute Zero (-273°C or 0 K)
The temperature at which all molecules stop moving.

Phase Change Diagrams

A **phase change diagram** shows the relationship between temperature and phase changes of a substance over a period of time as the substance is heating or cooling.

A **heating curve** shows a change of a substance starting with the substance as a solid. Changes represented on a heating curve are endothermic (heat is absorbed).

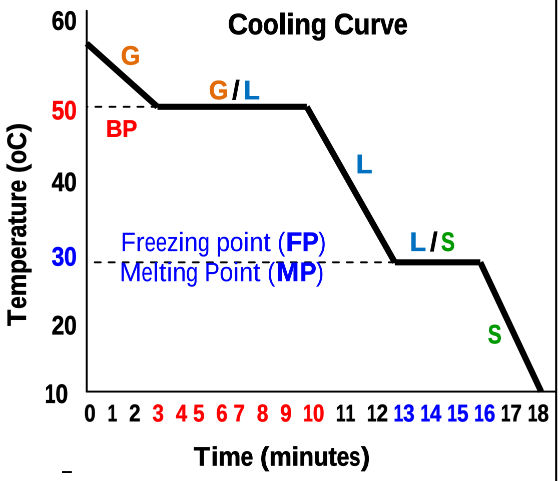
A **cooling curve** shows a change of a substance starting with the substance as a gas. Changes represented on a cooling curve are exothermic (heat is released).



During segments S, L, G.
.One phase is present
. Temperature increases
. Kinetic energy increases
. Potential energy stays same

During segments
S / L and L / G
Two phases are present
. Temperature stays same
. Kinetic energy stays same
. Potential energy increases

The substance represented by this curve is likely water.



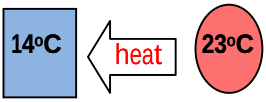
During segments G, L, S.
.One phase is present
. Temperature decreases
. Kinetic energy decreases
. Potential energy stays same

During segments
S / L and L / G
. Two phases are present
. Temperature stays same
. Kinetic energy stays same
. Potential energy decreases

The substance represented by this curve is not water.

Lesson 3: Heat and heat calculations

Introduction



Heat is a form of energy that can flow (or transfer) from one object to another. Direction of heat flow depends on the temperature difference. ***Heat flows*** from an area or object of a higher temperature to an area or object of a lower temperature until equilibrium temperature is reached. The equilibrium temperature in the above diagram will be 18.5^{o}C (The sum of the two temperatures divided by 2).

During chemical and physical changes, heat energy is either absorbed or released.

Exothermic describes a process that releases (emits or loses) heat.

Endothermic describes a process that absorbs (or gains) heat.

Joules and **calories** are the two most common units for measuring heat. A **calorimeter** is a device that is used for measuring heat during physical and chemical changes.

Heat constants and heat equations

Specific heat capacity (C) of a substance is the amount of heat needed to change the temperature of a one gram sample of the substance by just one degree Celsius (or one Kelvin). Specific heat capacity is different for different substances.

Specific heat capacity (C) for water is 4.18 J/g.^{o}K (See Table B). In other words, a one gram sample of water will absorb 4.18 Joules of heat to increase its temperature by one Kelvin, or release 4.18 Joules of heat to decrease its temperature by one Kelvin.

When the mass and specific heat capacity of a substance are known, the amount of heat absorbed or released by that substance to change between any two temperatures can be calculated using the **Table T** equation below.

$$\textbf{Heat (q)} = \textbf{m} \times \textbf{C} \times \Delta \textbf{T}$$

m = mass of substance (g)

C = specific heat capacity (J/g.K)

ΔT = difference in temp (K or ^{o}C)

ΔT = High temp - Low temp

Example:

How much heat is released by a 7 gram sample of water to change its temperature from 15 ^{o}C to 10 ^{o}C?

q = 7 x 4.18 x 5 *setup*

q = $\boxed{\textbf{146.3 J}}$ *calculated result*

Heat of fusion (Hf) of a substance is the amount of heat needed to melt a one gram sample of the substance at constant temperature. The heat of fusion for water is 334 J/g **(See Table B**). In other words, a one gram sample of water will absorb 334 Joules of heat to melt, or release 334 Joules of heat to freeze.

When the mass and heat of fusion of a substance are known, the amount of heat absorbed or released by the substance to change between the solid and liquid phases can be calculated using the **Table T** equation below.

Heat (q) = m x Hf

m = mass of substance (g)
Hf = Heat of fusion (J/g)

What is the number of joules needed to melt a 16 g sample of ice to water at 0°C?

q = m x Hf
q = 16 x 334 *setup*
q = $\boxed{\textbf{5344 J}}$

Heat of vaporization (Hv) of a substance is the amount of heat needed to vaporize (evaporate) a one gram sample of the substance at a constant temperature.

The heat of vaporization of water is 2260 J/g. In other words, a one gram sample of water will absorb 2260 Joules of heat to vaporize, or release 2260 Joules of heat to condense .

When the mass and heat of vaporization of a substance are known, the amount of heat absorbed or released by the substance to change between the liquid and gas phases can be calculated using **Table T** equation below:

Heat = m x Hv

m = mass of substance (g)
Hv = Heat of vaporization (J/g)

Liquid ammonia has a heat of vaporization of 1.35 kJ/g. How many kilojoules of heat are needed to evaporate a 5 gram sample of ammonia at its boiling point?

q = m x Hv
q = 5 x 1.35 *setup*
q = $\boxed{\textbf{6.75 kJ}}$

Solving a heat problem correctly depends on your understanding of the question, as well as choosing the right heat equation and substituting the correct factors into the equation. Keep the following key words or phrases in mind when deciding which of the three heat equations on Table T to choose.

Two temperatures given, changes temperature from: ***Choose q* = mCΔT**

To melt, to freeze, changes from liquid to solid, at 0°C : ***Choose q* = mHf**

To boil, to condense, changes from liquid to steam, at 100°C: ***Choose q* = mHv**

Lesson 4: Gas characteristics and gas laws

Introduction

Behavior of gases is influenced by three key factors: volume (space of container), pressure and temperature. The relationships between these three factors are the basis for gas laws and gas theories. These laws and theories attempt to explain how gases behave.
In this lesson you will learn about the gas laws and theories as well as gas law calculations.

Kinetic Molecular Theory of Ideal Gas

The **kinetic molecular theory** of an ideal gas is a model that is often used to explain the behavior of gases. This theory is summarized below.

- Gases are composed of individual particles
- Distances between gas particles are far apart
- Gas particles are in continuous, random, straight-line motion
- When two particles of a gas collide, energy is transferred from one particle to another
- Particles of a gas have no attraction to each other
- Individual gas particles have no volume (negligible or insignificant)

An **ideal gas** is a theoretical (or assumed) gas that has all properties summarized above.

A **real gas** is a gas that actually does exist. Examples of real gases are *oxygen, carbon dioxide, hydrogen, helium…etc..*

Since kinetic molecular theory (summarized above) applies mainly to an ideal gas, the model cannot be used to predict exact behavior of real gases. Therefore, real gases deviate from (do not behave exactly as) an ideal gas for the following reasons.

- Real gas particles do attract each other.
  Ideal gas particles are assumed to have no attraction to each other
- Real gas particles do have volume
  Ideal gases are assumed to have no volume.

Real gases with small molecular masses behave most like an ideal gas. Hydrogen (H) and Helium (He), the two smallest real gases by mass, will behave more like an ideal gas than any other real gas.

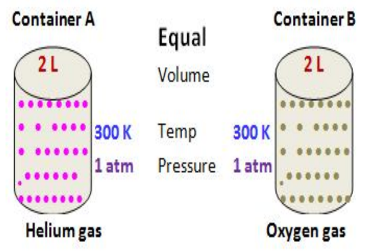
Real gases behave more like an ideal gas under conditions of
High temperature and Low pressure.

Helium, a real gas, will behave more like an ideal gas at
300 K and 1 atm. than at 273 K and 2 atm.

Gas laws

Avogadro's law states that: Under the same conditions of temperature and pressure, gases of **equal volume** contain **equal number of molecules (particles).**

Containers A and B to the right contain the same number of molecules.



Dalton's Law of Partial Pressure states: The total pressure (P_{total}) of a gas mixture is the sum of all the partial pressures.

Partial Pressure (P) is a pressure exerted by individual gas in a gas mixture

Total Pressure from Partial Pressures:

$$P_{total} = P_{gasA} + P_{gasB} + P_{gas\ C}$$

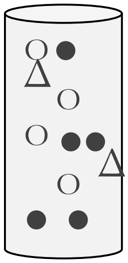
Total Pressure when gas X is collected over water:

$$P_{total} = P_{gas\ X} + VP_{H_2O} \text{ (at temp)}$$

VP_{H_2O} is the vapor pressure of water at the given water temperature. See **Table H** for vapor pressure at different temperatures.

Partial Pressure of gas X from mole fraction:

$$P_{gas\ X} = \frac{\text{moles of gas X}}{\text{total moles}} (P_{total})$$



a three-gas mixture

$P_{gas}\ \Delta = .2$ atm

$P_{gas}\ \text{O} = .4$ atm

$P_{gas}\ \bullet = .5$ atm

P_{total} = .2 + .4 + .5 = $\boxed{1.1 \text{ atm}}$

Oxygen gas is collected over water at 45°C in a test tube. If the total pressure of the gas mixture in the test tube is 26 kPa, what is the partial pressure of the oxygen gas ?

26 kPa = Pgas O + VP_{H_2O} at 45°C

26 kPa = Pgas O + **10**

$\boxed{16 \text{ kPa}}$ = Pgas O

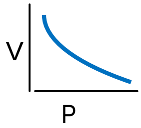
A gas mixture contains 0.8 moles of O_2 and 1.2 moles of N_2. If the total pressure of the mixture is 0.5 atm, what is the partial pressure of N_2 in this mixture?

$$P_{gas}\ N_2 = \frac{1.2}{2.0} \times 0.5 = \boxed{0.3 \text{ atm}}$$

Graham's law of Diffusion states that: The rate of diffusion (movement or spread) of a gas is proportional to its mass. I other words, a lighter gas will diffuse faster than a heavier gas.

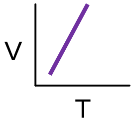
Boyle's Law states that: At **constant temperature**, volume of a gas is inversely proportional to the pressure on the gas. In other words, as pressure increases, volume (space) of the gas decreases by the same factor. The Boyle's law equation given below can be used to calculate the new volume of a gas when pressure on the gas is changed at constant temperature.

$$\boxed{P_1 V_1 = P_2 V_2}$$



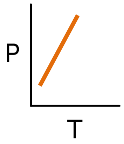
Charles's Law states that: At **constant pressure**, the volume of a gas is directly proportional to the *Kelvin* temperature of the gas. In other words, as temperature increases, volume (space) increases by the same factor. The Charles's law equation given below can be used to calculate the new volume of a gas when temperature of the gas is changed at constant pressure.

$$\boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$$



Gay-Lussac's Law states that: At **constant volume**, pressure of a gas is directly proportional to the *Kelvin* temperature of the gas. In other words, as temperature increases, pressure increases by the same factor.. The Gay-Lussac's law equation given below can be used to calculate the new pressure of a gas when temperature of the gas is changed at constant volume.

$$\boxed{\frac{P_1}{T_1} = \frac{P_2}{T_2}}$$



Combined gas law describes a gas behavior when all three factors (volume, pressure, and temperature) of the gas are changing: In the combined gas law, the only constant is the mass of the gas. The combined gas law equation below can be used to solve any problem related to the above three gas laws.

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

Table T equation

P = pressure
V = volume
T = Kelvin temperature
1 = initial condition
2 = new condition

At constant temperature, what is the new volume of a 3 L sample of O gas if its pressure is changed from 0.5 atm to 0.25 atm?

Setup

$(0.5)(3) = (0.25)(V_2)$

$\boxed{\textbf{6 L}} = \mathbf{V_2}$

The volume of a confined gas is 25 ml at 280 K. At what temperature would the gas volume be 75 ml if the pressure is held constant?

Setup

$$\frac{25}{280} = \frac{75}{T_2}$$

$\boxed{\textbf{840 K}} = \mathbf{T_2}$

At constant volume, pressure on a gas changes from 45 kPa to 50 kPa when the temperature of the gas is changed to 340K . What was the initial temperature of the gas?

Setup

$$\frac{45}{T_1} = \frac{50}{340}$$

$\mathbf{T_1} = \boxed{\textbf{306 K}}$

A 30 mL sample of H_2 gas is at 1 atm and 200 K. What will be its new volume at 2.0 atm and 600 K?

Setup

$$\frac{(1)(30)}{200} = \frac{(2.0)\,V_2}{600}$$

$\boxed{\textbf{45 mL}} = V_2$

Pressure, Volume, and Temperature

Pressure

Pressure of a gas is a measure of how much force is put on a confined gas.

Units: atmosphere (atm) or Kilopascal (kPa) ***1 atm = 101.3 kPa***

Volume

Volume of a gas measures the space a confined gas occupies (takes up).
Volume of a gas is the space of the container the gas is placed.

Units: milliliters (ml) or centimeters cube(cm^3) ***1 ml = 1 cm^3***
1 L = 1000 ml

Temperature

Temperature of a gas is a measure of the average kinetic energy of the gas particles. As temperature increases, gas particles move faster, and their average kinetic energy increases.

Units: degree Celsius (°C) or Kelvin (K) ***K = °C + 273***

Standard Temperature and Pressure: STP

Standard Temperature: **273 K** or **0°C**
Standard Pressure: **1 atm** or **101.3 kPa**

REFERENCE TABLE A

In some gas law problems, the temperature and/or pressure of the gas may be given at STP.

When a gas is said to be at STP in a gas law problem, the above values should be substituted into a gas law equation as needed. Be sure the unit of STP you choose is the same as the other unit in the given question.

NOTE: Always use Kelvin temperature in all gas law calculations.

Example;

Hydrogen gas has a volume of 100 mL at ***STP****. If temperature and pressure are changed to 546 K and 0.5 atm respectively, what will be the new volume of the gas?*

V_1 = 100 mL V_2 = ?
STP { **T_1 = 273 K** T_2 = 546 K
P_1 = 1 atm P_2 = 0.5 atm

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

$$\frac{(1)\,(100)}{273} = \frac{(0.5)\,(V_2)}{546}$$ *setup*

$$\boxed{\mathbf{400\ mL}} = \mathbf{V_2}$$ *calculated result*

Lesson 5: Physical and chemical properties and changes

Introduction

Properties are sets of characteristics that can be used to identify and classify matter. Two types of properties of matter are physical and chemical properties.

In this lesson, you will learn the differences between physical and chemical properties as well as the differences between physical and chemical changes of matter.

A **physical property** is a characteristic of a substance that can be observed or measured without changing the chemical composition of the substance. Some physical properties of a substance depend on sample size or amount, and some do not.

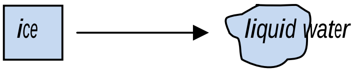
Extensive properties depend on sample size or amount present. Mass, weight and volume are examples of *extensive properties.*

Intensive properties do not depend on sample size or amount. Melting, freezing and boiling points, density, solubility, color, odor, conductivity, luster, and hardness are *intensive properties.*

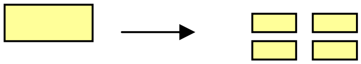
Differences in physical properties of substances make it possible to separate one substance from another in a mixture.

A **physical change** is a change of a substance from one form to another without changing its chemical composition.

Examples: Phase change



Size change



Dissolving NaCl(s) → Na+(aq) + Cl- (aq)

A **chemical property** is a characteristic of a substance that is observed or measured through interaction with other substances.

Examples:

It burns, it combusts, it decomposes, it reacts with, it combines with, or it rusts are some of the phrases that can be used to describe chemical properties of a substance.

A **chemical change** is a change in composition and properties of one substance to those of other substances. **Chemical reactions** are ways by which chemical changes of substances occur.

Types of chemical reactions include *synthesis, decomposition, single replacement, and double replacement*.

You will learn more about these reactions in Topic 5.

Practice Questions by Lessons

Lesson 1: Types of matter

Define the following terms and answer questions below.

1. Pure substance 2. Mixture 3. Element 4. Compound 5. Aqueous solution
6. Law of definite proportion 7. Homogeneous mixture
8. Heterogeneous mixture 9. Filtration 10. Distillation

11. Which property correctly describes all compounds?
1) They are always homogenous 3) They can be physically separated
2) They are always heterogeneous 4) They cannot be decomposed

12. Bronze contains 90 to 95 percent copper and 5 to 10 percent tin. Because these percentages can vary, bronze is classified as
1) A compound 2) A mixture 3) An element 4) A substance

13. When sample X is passed through a filter a white residue, Y, remains on the filter paper and a clear liquid, Z, passes through. When liquid Z is vaporized, another white residue remains. Sample X is best classified as
1) An element 3) A compound
2) A heterogeneous mixture 4) A homogeneous mixture

14. Which is a formula of a mixture of substances?
1) $Cl_2(g)$ 2) $MgCl_2(s)$ 3) $H_2O(l)$ 4) HF(aq)

15. The formula $N_2(g)$ is best classified as
1) A compound 2) A mixture 3) An element 4) A solution

Lesson 2: Phases of matter

Define the following terms and answer questions below.

16. Solid 17. Liquid 18. Gas 19. Condensation 20. Evaporation 21. Sublimation
22. Deposition 23. Exothermic 24. Endothermic 25. Temperature
26. Kinetic energy 27. Potential energy 28. Ice/liquid equilibrium
29. Water/steam equilibrium 30. Phase change diagram 31. Absolute Zero

32. Particles in which phase are arranged in a regular geometric pattern?
1) Solid 2) Aqueous 3) Liquid 4) Gas

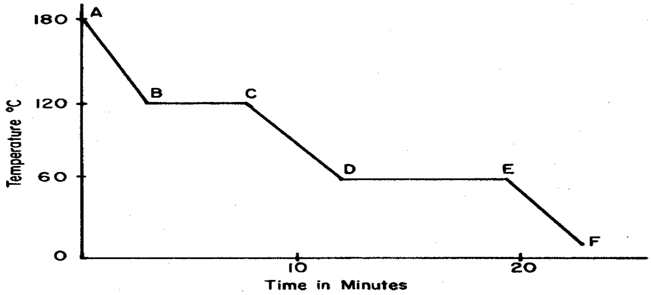
33. Which formula correctly represents a substance that has a definite volume but no definite shape?
1) Hg(l) 2) HCl(g) 3) Na(s) 4) $H_2(g)$

34. Which equation is showing sublimation of iodine?
1) $I_2(g) \text{-------}> I_2(s)$ 3) $I_2(s) \text{-------}> I_2(l)$
2) $I_2(s) \text{------}> I_2(g)$ 4) $I_2(g) \text{--------}> I_2(l)$

35. Which temperature of a solid substance will have particles with the highest kinetic energy?
1) 273 K 2) 373 K 3) 170°C 4) 70°C

36. Which change in temperature of a sample of water would result in the smallest decrease in the average kinetic energy of its molecules?
1) 25°C to 32°C 3) 15°C to 9°C
2) 25°C to 29°C 4) 12°C to 2°C

Answer questions 37 and 38 based on the information and diagram below.
The graph below represents the uniform cooling of an unknown substance, starting with the substance as a gas above its boiling point.



37. What is the melting point of the substance?
1) 0°C 2) 60°C 3) 120°C 4) 180°C

38. During which segment is the substance's kinetic energy remaining constant?
1) AB 2) BC 4) CD 4) EF

Lesson 3: Heat and heat calculations
Define the following terms and answer multiple choice questions below.
39. Heat 40. Joules 41. Specific heat capacity 42. Heat of fusion
43. Heat of vaporization 44. Calorimeter

45. The heat of fusion of ice is 334 Joules per gram. Adding 334 Joules to one gram of ice at STP will cause the ice to
1) Increase in temperature
2) Decrease in Temperature
3) Change to water at a higher temperature
4) Change to water at the same temperature

46. A solid material X is place in liquid Y. Heat will flow from Y to X when the temperature of
1) Y is 20°C and X is 30°C 3) Y is 15°C and X 10°C
2) Y is 10°C and X is 20°C 4) Y is 30°C and X is 40°C

47. How many kilojoules of heat are needed to raise the temperature of 500 g of water from 15°C to 20°C?
1) 4.20 KJ 2) 10.5 KJ 3) 32.0 KJ 4) 105 KJ

48. What amount of heat energy is needed to change a 20 g sample of water at 100°C to steam at the same temperature?
1) 905 KJ 2) 0.200 KJ 3) 1.13 KJ 4) 45.2 KJ

49. What is the total number of joules of heat energy released by a 2.5 gram sample of water to change to ice at 0°C?
1) 133 J 2) 8.4 J 3) 10.5 J 4) 835 J

50. What is the heat of vaporization of an unknown liquid if 5 grams of this liquid requires 22 KJ of heat to change to vapor at its boiling point?
1) 4.4 J/g 2) 100 J/g 3) 4400 J/g 4) 11300 J/g

Lesson 4: Gas laws and gas law calculations

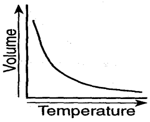
Define the following terms and answer multiple choice questions below.

51. Ideal gas 52.Kinetic molecular theory 53. Avogadro's law 54. Boyle's law
55.Charles' law 56. Gay-Lussac's law 57. Dalton's law of partial pressure

58. The kinetic molecular theory assumes that the particles of an ideal gas
1) Are in random, constant, straight-line motion
2) Are arranged in a regular geometric pattern
3) Have strong attractive forces between them
4) Have collisions that result in the system losing energy

59. Under which two conditions do real gases behave least like an ideal gas?
1) High pressure and low temperature 3) High pressure and high temperature
2) Low pressure and high temperature 4) Low pressure and low temperature

60. Which graph best illustrates the relationship between the Kelvin temperature of a gas and its volume when the pressure on the gas is held constant?

1)


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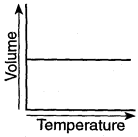

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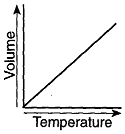

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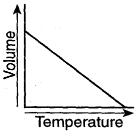


61. Which gas is least likely to obey the ideal gas model under the same temperature and pressure?
1) Xe 2) Kr 3) Ne 4) He

62. A real gas will behave most like an ideal gas under which conditions of temperature and pressure?
1) 273 K and 1 atm 3) 546 K and 2 atm
2) 273 K and 2 atm 4) 546 K and 1 atm

63. Under which conditions would a 2 L sample of O_2 has the same number of molecules as a 2 L sample of N_2 that is at STP?
1) 0 K and 1 atm 3) 0 K and 2 atm
2) 273 K and 1 atm 4) 273 K and 2 atm

64. A gas sample has a volume of 12 liters at 0^{o}C and 0.5 atm. What will be the new volume of the gas when the pressure is changed to 1 atm and the temperature is held constant?
1) 24 L 2) 18 L 3) 12 L 4) 6.0 L

65. At STP, a gas has a volume of 250 mL. If the pressure remained constant, at what Kelvin temperature would the gas have a volume of 50 ml?
1) 137 K 2) 500 K 3) 546 K 4) 273 K

66. A gas has a pressure of 120 kPa and a volume of 50.0 milliliters when its temperature is 127^{o}C. What volume will the gas occupy at a pressure of 60 kPa and at a temperature of -73^{o}C?
1) 12.5 ml 2) 50.0 ml 3) 100 ml 4) 200 ml

Lesson 5: Physical and chemical properties and changes

Define the following terms and answer multiple choice questions below.

67. physical property
68. chemical property
69. physical change
70. chemical change

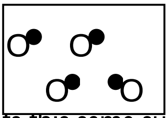
71. Which is a physical property of sodium?
    1) It is flammable
    2) It is shiny
    3) It reacts with water
    4) It reacts with chlorine

72. Which is a chemical property of water?
    1) It freezes
    2) It evaporates
    3) It boils
    4) It decomposes

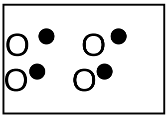
73. Which is a physical change of iodine?
    1) It can decompose into two iodine atoms
    2) Iodine can dissolve in water
    3) Iodine can react with sugar
    4) Iodine can react with hydrogen

74. An example of a chemical change is
    1) Boiling of water
    2) Dissolving of sodium bromide
    3) Burning of magnesium
    4) Breaking of sulfur into pieces

75. Given the particle diagram representing four molecules of a substance.



Which particle diagram best represents this same substance after a physical change has taken place?

1)


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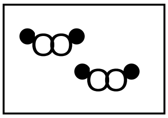

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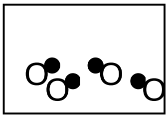

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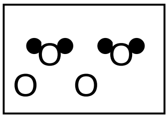
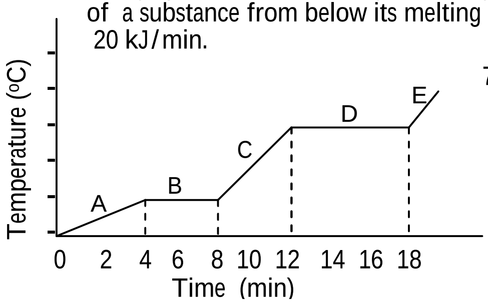


Topic Mastery

76. A 12 gram sample of water initially at 32^{o}C loses 780 joules of heat. What is the new temperature of the water?

77. A 50 L sample of O_2 gas is at STP. When the temperature of the gas is changed to 64^{o}C, the new volume of the gas is 20 L. What is the new pressure of the gas in kilopascals?

The graph below is showing the change in temperature of a 15 g sample of a substance from below its melting point as heat is added at a rate of 20 kJ/min.



78. Calculate the heat of fusion of the substance.

79. a) How would the heat of vaporization of the substance compares to the heat of fusion?

    b) Explain your answer using information from the graph.

Lesson 1: Arrangement of the Elements

Introduction

There are more than 100 known elements. Most of the elements are naturally occurring, while a few are artificially produced. The modern Periodic Table contains all known elements. These elements are arranged on the Periodic Table in the order of increasing atomic number.

Important information about an element can be found in the box of the element on the Periodic Table .

In this lesson, you will learn about the arrangement of the elements on the Periodic Table.

Properties of the Modern Periodic Table

The modern Periodic Table, which was created by Dmitri Mendeleev, has the following properties:

- Elements are arranged in the order of increasing atomic number
- The three types of elements found on the Periodic Table are metals, nonmetals, and metalloids
- More than two thirds (majority) of the elements are metals
- The Periodic Table contains elements that are in all three phases (solid, liquid, and gas) at STP
- The majority of the elements exist as solids
- Only two (mercury and bromine) are liquids. A few are gases.
- An element's symbol can be one (O), two (Na), or three (Uub) letters. The first letter must always be capitalized. The Second (or third) letter must be lowercase.

15.999	**Atomic Mass**	**196.967**
-2	*Selected Oxidation states (charges)*	*+1* *+2*
O	**Element's symbol**	**Au**
8	Atomic number	79
2 - 6	Electron configuration	2 - 8 -18 - 32 -18 - 1

Information listed in the box for each element is related to the atomic structure of that element. The Atomic Structure is discussed in Topic 3.

Groups and Periods

Groups are the vertical arrangements of the elements. There are 18 groups on the Periodic Table of the Elements. Group names are listed below.

Group 1 :	Alkali metals
Group 2 :	Alkaline earth metals
Group 3 – 12:	Transition metals
Group 17:	Halogens
Group 18:	Noble (Inert) gases

Elements in the same group have the same number of valence electrons. Valence electrons are electrons in the outermost energy level of an atom. Elements in the same group have similar chemical properties and reactivity due to similarity in their number of valence electrons.

Periods are the horizontal rows of the Periodic Table. Elements in the same period have the same number of occupied electron shells. There are seven (7) Periods on the Periodic Table of the Elements.

Periodic Law states that: The properties of the elements are a periodic function of their atomic numbers. In other words, by arranging the elements in order of increasing atomic number, a new period of elements is formed so that elements with similar chemical properties fall in the same group.

Allotropes

Allotropes are different molecular forms of the same element in the same state.

Allotropes of the same element have different molecular structures.

Differences in molecular structures give allotropes of the same element different physical properties (color, shape, and density, mass..) AND different chemical properties (reactivity).

Examples of some common allotropes:

Oxygen allotropes: Air (O_2) and Ozone (O_3)

Carbon allotropes: Diamond, graphite, and buckminsterfullerene

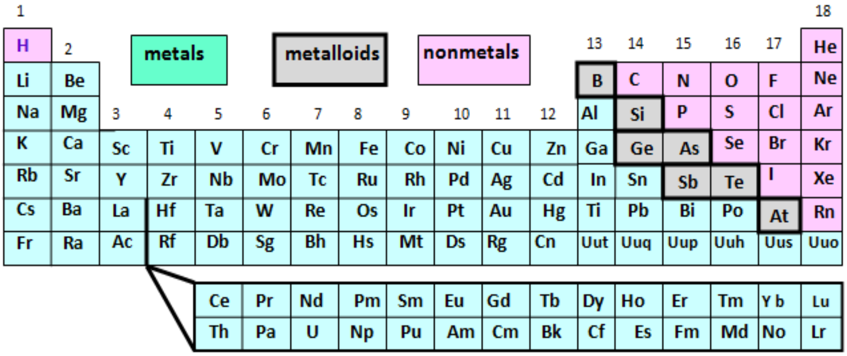
Phosphorous allotropes: Red, Black, and White

Lesson 2: Types of elements and their Properties

Introduction

There are three general categories of elements: metals, nonmetals and metalloids. Elements in each category have a set of physical and chemical properties that can be used to distinguish them apart from elements in other categories.
In this lesson, you will learn about the three different types of elements, their location on the Periodic Table, and their properties.

Location of metals, metalloids, and nonmetals



metals

metalloids

nonmetals

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Ti	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Uut	Uuq	Uup	Uuh	Uus	Uuo
				Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
				Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Properties of the Elements

There are several physical properties that are used to describe and identify the elements. Below are terms and definitions of these properties.

Malleable describes a solid that is easily hammered into a thin sheet.

Ductile describes a solid that is easily drawn into thin wire.

Brittle describes a solid that is easily broken or shattered into pieces when struck

Luster describes the shininess of a substance.

Conductivity describes the ability to conduct heat or electricity.

Electronegativity describes an atom's ability to attract electrons from another atom during bonding.

Ionization energy describes an atom's ability to lose its most loosely bound valence electrons.

Atomic radius describes the size of the atom of an element.

Density describes the mass to volume ratio of an element

See Table S for values to these four properties

Ionic radius describes the size of the element after it has lost or gained electrons to form an ion.

Properties of metals, metalloids, and nonmetals

Metal elements are located to the left of the Periodic Table.
All elements in Group 1 – 12 (except hydrogen) are classified as a metal. The rest of the metal elements are found near the bottom of Groups 13, 14 and 15. The majority (about 75%) of the elements are metals.

General properties of metals are listed below.

. All metals (except Hg) exist as a solid at STP. Hg is the only liquid metal.
. Metals are malleable, ductile, and have luster
. Metals tend to have high conductivity due to their mobile valence electrons
. Metals tend to have low electronegativity values (because they do not attract electrons easily)
. Metals tend to have low ionization energy values (because they lose their electrons easily)
. Metals lose electrons and form a positive ion during chemical bonding
. Radius (size) of a metal atom decreases as it loses electrons and forms a positive (+) ion
. The size of a +metal ion (ionic radius) is always smaller than the size of the neutral atom (atomic radius)

Metalloids are the elements located between the metals and the nonmetals. Metalloid elements are located along the zigzag line of the Periodic Table.

General properties of metalloids are listed below.

. Metalloids tend to have properties of both the metals and nonmetals
. Metalloid properties are more like those of metals and less like nonmetals
. Metalloids exist only as solids at STP

Nonmetal elements are located to the right of the Periodic Table.
All elements in Groups 17 and 18 are classified as nonmetals. The rest of the nonmetals are found near the top of Groups 14, 15, and 16. Hydrogen is also a nonmetal.

General properties of nonmetals are listed below.

. Nonmetals are found in all three phases: solid, liquid, and gas.
. Most nonmetals are either a gas or solid at STP. Br is the only liquid nonmetal
. Solid nonmetals are generally brittle and dull (lack luster, not shiny)
. Nonmetals have low (poor) electrical and heat conductivity
. Nonmetals tend to have high electronegativity values (because they attract or gain electrons easily)
. Nonmetals tend to have high ionization values (because they do not lose their electrons easily)
. Nonmetals generally gain electrons and form a negative ion during bonding
. Radius of a nonmetal atom increases as it gains electrons and forms a negative (–) ion
. The size of the - nonmetal ion (ionic radius) is always bigger than that of the neutral atom (atomic radius)

Summary of properties

	Phases at STP	***Physical properties***	***Conductivity***	***Electronegativity***	***Ionization energy***	***In bonding***	***Common ion***	***Ionic size (radius)***
Metals	Solid Liquid	Malleable Luster Ductile	High	Low	Low	Lose electrons	+ (positive)	Smaller than atom
Nonmetals	Solid Liquid Gas	Brittle Dull	Low	High	High	Gain electron	**-** (negative)	Bigger than atom
Metalloids	Solid only	Properties of metals and nonmetals	Low	-	-	Lose electrons	+ (positive)	Smaller than atom

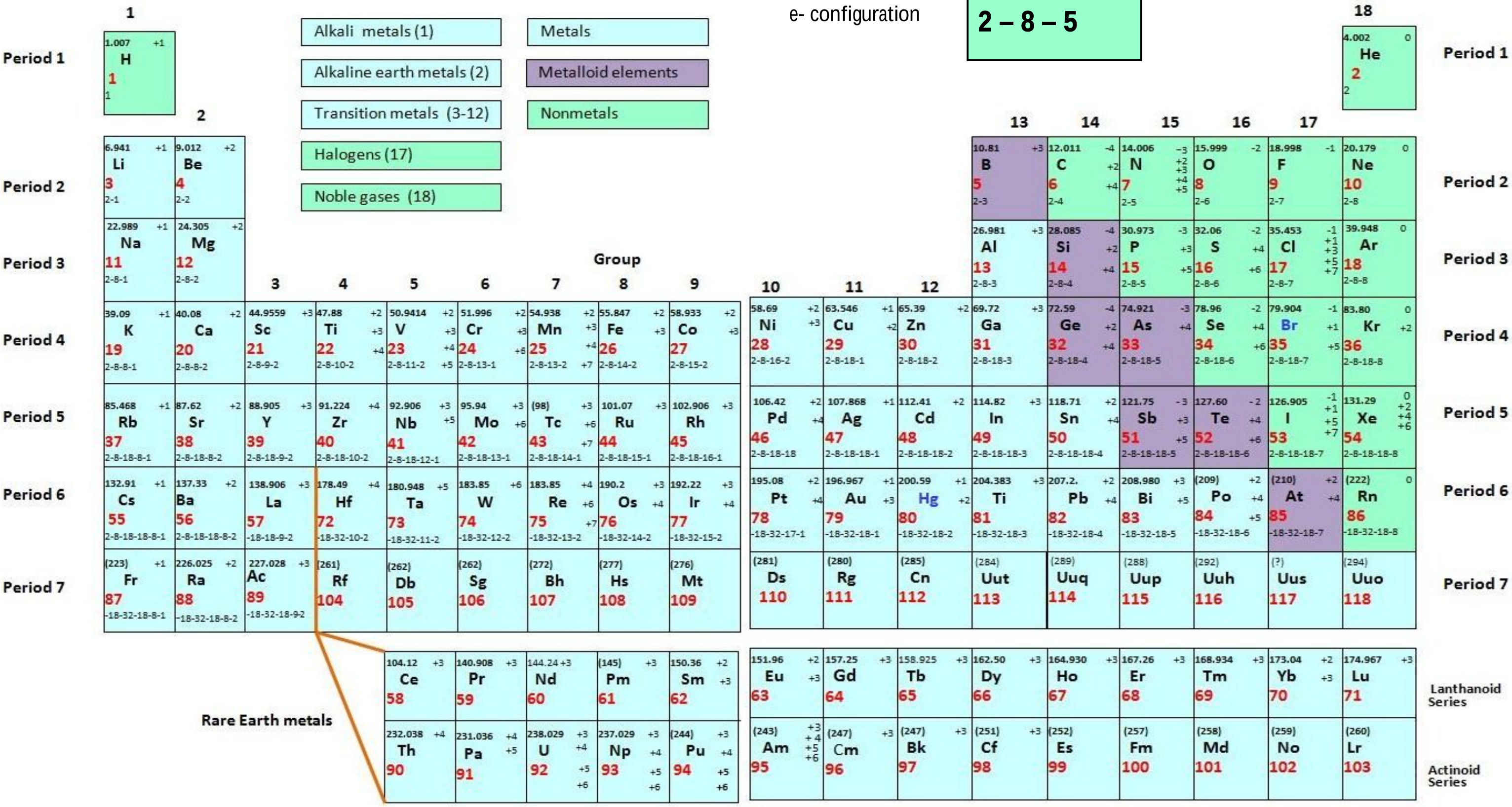
Properties of Groups

According to the Periodic Law, an element falls into a particular group based on its properties. Elements with similar chemical properties belong in the same group.

Below is a table summarizing group names and general characteristics of each group.

Group number	*Group name*	*Types of elements in the group*	*Phases (at STP)*	*Valance electrons (during bonding)*	*Common oxidation number (charge)*	*Chemical bonding (general formula)*
1	Alkali metals	Metal	Solid (all)	1 (lose)	+1	**XY** with halogens (17) **X_2O** with oxygen (16)
2	Alkaline earth	Metal	Solid (all)	2 (lose)	+2	**MY_2** with halogens (17) **MO** with oxygen (16))
3-12	Transition metals	Metal	Liquid (Hg) Solid (the rest)	(lose)	Multiple + charges	varies (form colorful compounds)
13	-	Metalloid Metal	Solid (all	3 (lose)	+3	**LY_3** with halogens (17) **L_2O_3** with oxygen (16)
14	-	Nonmetal Metalloid Metal	Solid (all)	4 (some share) (some lose)	vary	varies
15	-	Nonmetal Metalloid Metal	Gas (N) Solid (the rest)	5 (gain or share)	-3	varies
16	Oxygen group	Nonmetal Metalloid	Gas (O) Solid (the rest)	6 (gain or share)	-2	**X_2O** with alkali metals (1) **MO** with alkaline earth (2)
17	Halogens (Diatomic)	Nonmetal	Gas (F and Cl) Liquid (Br) Solid (I)	7 (gain or share)	-1	**XY** with alkali metals (1) **MY** with alkaline earths (2)
18	Noble gases (Monatomic)	Nonmetal	Gas (all)	8 (neither gain nor share)	0	Forms very few compounds. **XeF_4** is the most common.

The Periodic Table of the Elements



Alkali metals (1)

Alkaline earth metals (2)

Transition metals (3-12)

Halogens (17)

Noble gases (18)

Metals

Metalloid elements

Nonmetals

Atomic Mass	30.973	Selected oxidation states -3 +3 +5
Element's Symbol	P	
Atomic Number	15	
e- configuration	2 – 8 – 5	

Group

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	
Period 1	1.007 +1 H 1 1																	4.002 0 He 2 2	Period 1
Period 2	6.941 +1 Li 3 2-1	9.012 +2 Be 4 2-2											10.81 +3 B 5 2-3	12.011 -4 +2 +4 C 6 2-4	14.006 -3 +2 +3 +4 +5 N 7 2-5	15.999 -2 O 8 2-6	18.998 -1 F 9 2-7	20.179 0 Ne 10 2-8	Period 2
Period 3	22.989 +1 Na 11 2-8-1	24.305 +2 Mg 12 2-8-2											26.981 +3 Al 13 2-8-3	28.085 -4 +2 +4 Si 14 2-8-4	30.973 -3 +3 +5 P 15 2-8-5	32.06 -2 +4 +6 S 16 2-8-6	35.453 -1 +1 +3 +5 +7 Cl 17 2-8-7	39.948 0 Ar 18 2-8-8	Period 3
Period 4	39.09 +1 K 19 2-8-8-1	40.08 +2 Ca 20 2-8-8-2	44.9559 +3 Sc 21 2-8-9-2	47.88 +2 +3 +4 Ti 22 2-8-10-2	50.9414 +2 +3 +4 +5 V 23 2-8-11-2	51.996 +2 +3 +6 Cr 24 2-8-13-1	54.938 +2 +3 +4 +7 Mn 25 2-8-13-2	55.847 +2 +3 Fe 26 2-8-14-2	58.933 +2 +3 Co 27 2-8-15-2	58.69 +2 +3 Ni 28 2-8-16-2	63.546 +1 +2 Cu 29 2-8-18-1	65.39 +2 Zn 30 2-8-18-2	69.72 +3 Ga 31 2-8-18-3	72.59 -4 +2 +4 Ge 32 2-8-18-4	74.921 -3 +4 As 33 2-8-18-5	78.96 -2 +4 +6 Se 34 2-8-18-6	79.904 -1 +1 +5 Br 35 2-8-18-7	83.80 0 +2 Kr 36 2-8-18-8	Period 4
Period 5	85.468 +1 Rb 37 2-8-18-8-1	87.62 +2 Sr 38 2-8-18-8-2	88.905 +3 Y 39 2-8-18-9-2	91.224 +4 Zr 40 2-8-18-10-2	92.906 +3 +5 Nb 41 2-8-18-12-1	95.94 +3 +6 Mo 42 2-8-18-13-1	(98) +3 +6 +7 Tc 43 2-8-18-14-1	101.07 +3 Ru 44 2-8-18-15-1	102.906 +3 Rh 45 2-8-18-16-1	106.42 +2 +4 Pd 46 2-8-18-18	107.868 +1 Ag 47 2-8-18-18-1	112.41 +2 Cd 48 2-8-18-18-2	114.82 +3 In 49 2-8-18-18-3	118.71 +2 +4 Sn 50 2-8-18-18-4	121.75 -3 +3 +5 Sb 51 2-8-18-18-5	127.60 -2 +4 +6 Te 52 2-8-18-18-6	126.905 -1 +1 +5 +7 I 53 2-8-18-18-7	131.29 0 +2 +4 +6 Xe 54 2-8-18-18-8	Period 5
Period 6	132.91 +1 Cs 55 2-8-18-18-8-1	137.33 +2 Ba 56 2-8-18-18-8-2	138.906 +3 La 57 -18-18-9-2	178.49 +4 Hf 72 -18-32-10-2	180.948 +5 Ta 73 -18-32-11-2	183.85 +6 W 74 -18-32-12-2	183.85 +4 +6 +7 Re 75 -18-32-13-2	190.2 +3 +4 Os 76 -18-32-14-2	192.22 +3 +4 Ir 77 -18-32-15-2	195.08 +2 +4 Pt 78 -18-32-17-1	196.967 +1 +3 Au 79 -18-32-18-1	200.59 +1 +2 Hg 80 -18-32-18-2	204.383 +3 Ti 81 -18-32-18-3	207.2. +2 +4 Pb 82 -18-32-18-4	208.980 +3 +5 Bi 83 -18-32-18-5	(209) +2 +4 +5 Po 84 -18-32-18-6	(210) +2 +4 At 85 -18-32-18-7	(222) 0 Rn 86 -18-32-18-8	Period 6
Period 7	(223) +1 Fr 87 -18-32-18-8-1	226.025 +2 Ra 88 -18-32-18-8-2	227.028 +3 Ac 89 -18-32-18-9-2	(261) Rf 104	(262) Db 105	(262) Sg 106	(272) Bh 107	(277) Hs 108	(276) Mt 109	(281) Ds 110	(280) Rg 111	(285) Cn 112	(284) Uut 113	(289) Uuq 114	(288) Uup 115	(292) Uuh 116	(?) Uus 117	(294) Uuo 118	Period 7

Rare Earth metals

104.12 +3 Ce 58	140.908 +3 Pr 59	144.24 +3 Nd 60	(145) +3 Pm 61	150.36 +2 +3 Sm 62	151.96 +2 +3 Eu 63	157.25 +3 Gd 64	158.925 +3 Tb 65	162.50 +3 Dy 66	164.930 +3 Ho 67	167.26 +3 Er 68	168.934 +3 Tm 69	173.04 +2 +3 Yb 70	174.967 +3 Lu 71	Lanthanoid Series
232.038 +4 Th 90	231.036 +4 +5 Pa 91	238.029 +3 +4 +5 +6 U 92	237.029 +3 +4 +5 +6 Np 93	(244) +3 +4 +5 +6 Pu 94	(243) +3 +4 +5 +6 Am 95	(247) +3 Cm 96	(247) +3 Bk 97	(251) +3 Cf 98	(252) Es 99	(257) Fm 100	(258) Md 101	(259) No 102	(260) Lr 103	Actinoid Series

Group Names and Characteristics (also see table on page 23)

Group 1: Alkali metals

. Found in nature as compounds (not as free elements) due to high reactivity

. Are obtained from electrolytic reduction of fused salts (NaCl, KBr ..etc)

. Francium is the most reactive metal in Group 1, and of all metals

. Francium is also radioactive

. All alkali metals exist as solids at room temperature

Group 2: Alkaline Earth metals

. Found in nature as compounds (not as free element) due to high reactivity.

. Are obtained from fused salt compounds ($MgCl_2$, $CaBr_2$..etc)

. All alkaline earth metals exist as solids at room temperature

Group 3 – 12: Transition metals

. Properties of these elements vary widely

. They tend to form multiple oxidation numbers

. Most can lose electrons in two or more different sublevels of their atoms

. Their ions usually form colorful compounds

Examples: **Cu**Cl_2 – is a bluish color compound
FeCl_2 - is a reddish-orange color compound

Group 17: Halogens

. Exist as diatomic (two-atom) molecules (F_2, Cl_2, Br_2)

. The only group with elements in all three phases at STP

. Fluorine is the most reactive of the group, and of all nonmetals

. Fluorine is obtained from fused salt compounds (NaF, NaCl..etc)

. Astatine (At) in this group is a metalloid

Group 18: Noble Gases

. Exist as monatomic (one-atom) molecules (Ne, He, Kr…)

. They all have full and stable valence shells with 8 electrons (He is full with just 2 electrons)

. All are very stable and non-reactive (do not form many compounds)

. Argon(Ar) and Xenon(Xe) have been found to produce a few stable compounds with fluorine.

Ex. **Xe**F_4 (xenon tetrafluoride)

Lesson 3: Periodic Trends

Introduction

Periodic trends refer to patterns of properties that exist as elements are considered from one end of the table to the other.

Trend in atomic number is a good example (and the most obvious) of a periodic trend found on the Periodic Table.

As elements are considered one after the other from:

Left to **Right** across a Period: Atomic number of the elements increases.

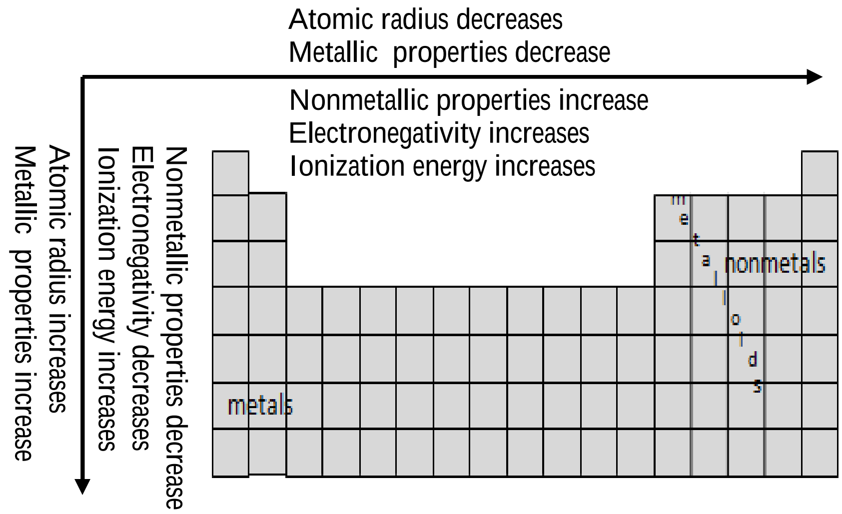
Bottom to **Top** up a Group: Atomic number of the elements decreases.

Many other trends exist on the Periodic Table even though they may not be so obvious.

In this lesson, you will learn of the following trends.

Trends in atomic and ionic radius (size).
Trends in metallic and nonmetallic properties.
Trends in electronegativity and ionization energy.

Summary of Periodic Trends



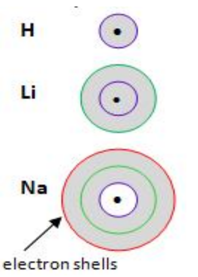
Trends in Atomic Radius

Atomic radius is defined as half the distance between two nuclei of the same atom when they are joined together.

Atomic radius measurement gives a good approximation of the size of each atom. The trend in atomic radius is as follows.

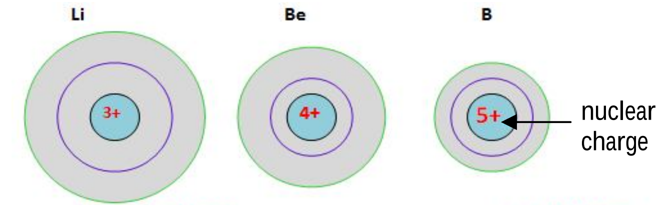
Top to Bottom down a Group:
Atomic size increases due to an increase in the number of ***electron shells.***

Left to Right across a Period:
Atomic size (radius) decreases due to an increase in ***nuclear charge.***



One shell:
Smallest radius

Three shells:
Largest radius



Smallest nuclear charge
Biggest radius (size)

Greatest nuclear charge
Smallest radius (size)

Use Reference Table S to note and compare atomic radii of atoms.

Trends in Metallic and Nonmetallic properties

Trends in properties and reactivity vary between metals and nonmetals. The bottom left corner contains the most reactive metals. ***Francium*** is the most reactive of all metals. The top right corner contains the most reactive nonmetals. ***Fluorine*** is the most reactive of all nonmetals.

Trends in metallic and nonmetallic properties and reactivity are summarized below.

Top to **Bottom** down a **Group**:
Metallic properties and reactivity increase *(ex. K is more reactive than Na)*
Nonmetallic properties and reactivity decrease *(ex. Br is less reactive than Cl)*

LEFT to **Right** across a **Period**:
Metallic properties and reactivity decrease. *(ex. Mg is less metallic than Na)*
Nonmetallic properties and reactivity increase. *(ex. Cl is more nonmetallic than S)*

Trends in Electronegativity and Ionization Energy

Electronegativity defines an atom's ability to attract (or gain) electrons from another atom during chemical bonding. The electronegativity value assigned to each element is relative to one another. The higher the electronegativity value, the more likely it is for the atom to attract (or gain) electrons and form a negative ion during bonding.
Fluorine (F) is assigned the highest electronegativity value of 4.
Francium (Fr) is assigned the lowest electronegativity value of 0.7 .
This means that of all the elements, fluorine has the greatest tendency to attract (or gain) electrons. Francium has the least ability or tendency to attract electrons during bonding.

Ionization energy refers to the amount of energy needed to remove an electron from an atom. The ***first ionization energy*** is the energy to remove the most loosely bound electron from an atom. Ionization energy measures the tendency of (how likely) an atom to lose electrons and form a positive ion. The lower the first ionization energy of an atom, the easier (the more likely) it is for that atom to lose its most loosely bound valence electron and form a positive ion.

Metals lose electrons because of their low ionization energies. The *alkali metals* in Group 1 generally have the lowest ionization energy, which allows them to lose their one valence electron most readily.

Nonmetals have low tendency to lose electrons because of their high ionization energies. The *noble gases* in group 18 tend to have the highest ionization energy values. Since these elements already have a full valence shell of electrons, a high amount of energy is required to remove any electron from their atoms.

Trends in electronegativity and ionization energy are as follows.

Top to **Bottom** down a **Group**:
Electronegativity (tendency to gain or attract electrons) decreases
due to increase in atomic sizes.
ex. S will attract electrons less readily than O because S is bigger than O

Ionization energy (tendency to lose or give up electrons) decreases
due to increase in atomic sizes.
ex. S will lose electrons more readily than O because S is bigger than O

Left to **Right** across a **Period**:
Electronegativity increases due to decrease in atomic sizes.
ex. S will attract electrons more readily than P because S is smaller than P

Ionization energy increases due to decrease in atomic sizes.
ex. S will lose electrons less readily than P because S is smaller than P

Use Reference Table S to note and compare electronegativity and ionization energy values of the elements.

Practice Questions by Lessons

Lesson 1: Arrangements of the elements

Define the following terms and answer multiple choice questions below.

1. Periodic Law 2. Group 3. Period 4. Allotrope

5. The observed regularities in the properties of the elements are periodic functions of their
   1) Oxidation state 2) Atomic numbers 3) Atomic mass 4) Reactivity

6. Which of the following information cannot be found in the box of elements on the Periodic Table?
   1) Oxidation state 2) Atomic number 3) Atomic mass 4) Phase

7. In general, elements within each group of the Periodic Table share similar
   1) Chemical properties 3) Mass number
   2) Electron configuration 4) Number of occupied energy levels

8. Which list contains elements with the greatest variation in chemical properties?
   1) O, S and Se 2) N, P and As 3) Be, N, O 4) Ba, Sr and Ca

9. Which element has similar chemical reactivity to the element chlorine?
   1) Bromine 2) Sulfur 3) Argon 4) Calcium

10. Oxygen and sulfur can both form a bond with sodium with similar chemical formula. The similarity in their formulas is due to
    1) Oxygen and sulfur having the same number of kernel electrons
    2) Oxygen and sulfur having the same number of valence electrons
    3) Oxygen and sulfur having the same number of protons
    4) Oxygen and sulfur having the same molecular structure

Lesson 2: Types of elements and properties

Define the followings terms and answer multiple choice questions below.

11. Malleable 12. Luster 13. Brittleness 14. Ductile 15. Ionization energy
16. Electronegativity 17. Density 18. Atomic radius 19. Alkali metal
20. Alkaline earth metal 21. Transition element 22. Halogen. 23. Noble gas

24. Solid nonmetal elements tend to be
    1) Malleable 2) Brittle 3) Ductile 4) Luster

25. An element has luster as one of its physical properties. Which is true of this element?
    1) It is a gas 2) It is a metal 3) It is a nonmetal 4) It is a halogen

26. Which properties are characteristics of metallic elements?
    1) Low ionization energy and malleable 3) Brittleness and dullness
    2) Low heat conductivity and luster 4) Brittleness and ductile

27. Which physical characteristic of a solution indicates the presence of a transition element?
    1) Its effect on litmus 2) Its density 3) Its color 4) Its reactivity

28. Element X is a solid at STP. Element X could be a
    1) Metal
    2) Nonmetal
    3) Metalloid
    4) Metal, nonmetal, or metalloid

29. Which element is a metalloid?
    1) B 2) Al 3) Sn 4) Au

30. Which group contains only metallic elements?
    1) Group 2 2) Group 13 3) Group 14 4) Group 17

31. Which of these elements in Period 2 is likely to form a negative ion?
    1) Oxygen 2) Boron 3) Ne 4) Li

32. Which properties best describe the element silver?
    1) Malleable and low electrical conductivity
    2) Brittle and low electrical conductivity
    3) Malleable and high electrical conductivity
    4) Brittle and high electrical conductivity

33. Which set contains elements that are never found in nature in their atomic state?
    1) C and Na 2) K and S 3) Na and P 4) Na and K

34. A Period 2 element forms a compound with oxygen with a formula of Z_2O? Element Z could be
    1) Neon 2) Boron 3) Be 4) Li

35. Element L is in Period 3 of the Periodic Table. Which element is L if it forms a compound with bromine with the formula LBr_3?
    1) Na 2) Mg 3) Al 4) Cl

36. Elements potassium and cesium are both classified as
    1) Transition metals 2) Alkali metals 3) Halogens 4) Noble gases

Lesson 3: Periodic Trends

Answer the following multiple choice questions.

37. As the elements in Group 1 of the Periodic Table are considered in order of increasing atomic number, the atomic radius of each successive element increases. This is primarily due to an increase in the number of
    1) Neutrons in the nucleus
    2) Unpaired electrons
    3) Valence electrons
    4) Electron shells

38. When the elements within Group 16 are considered in order of increasing atomic number, the electronegativity value of successive elements
    1) Increases 2) Decreases 3) Remains the same

39. When the elements within a period on the Periodic Table are considered in order of increasing atomic number, the nonmetallic properties of successive elements
    1) Increases 2) Decreases 3) Remains the same

40. When elements within Group 16 are considered in order of decreasing atomic number , the first ionization energy of successive elements generally
    1) Increases 2) Decreases 3) Remains the same

41. As the halogens in Group 17 are considered in order from bottom to top , the number of valence electrons of successive elements generally
    1) Increases 2) Decreases 3) Remains the same

42. Which of these Group 14 elements has the smallest atomic radius?
    1) Lead 2) Tin 3) Silicon 4) Carbon

43. Which atom has a bigger atomic radius than an atom of sulfur?
    1) Oxygen 2) Phosphorous 3) Chlorine 4) Argon

44. According to the Periodic Table, which sequence correctly places the elements in order of increasing atomic size?
    1) Na ---- > Li ---- > H ---- > K
    3) Te ---- > Sb ----- > Sn ---- > In
    2) Ba ---- > Sr ---- >Mg --- > Ca
    4) H ----- > He ----- > Li ---- > Be

45. Which of these halogens is the most reactive on the Period Table?
    1) I 2) Br 3) Cl 4) F

46. Which of these elements has the most metallic properties ?
    1) Radium 2) Strontium 3) Magnesium 4) Beryllium

47. Which element has the least tendency to lose its valence electrons during bonding?
    1) Potassium 2) Selenium 3) Bromine 4) Calcium

48. Which element has the greatest tendency to attract electrons during bonding?
    1) Se 2) S 3) Te 4) O

49. Which sequence of elements is arranged in order of decreasing tendency to attract electrons during chemical bonding?
    1) Al, Si, P 2) Cs, Na, Li 3) I, Br, Cl 4) C, B, Be

50. Which of these Group 2 elements has the highest eletronegativity value?
    1) Be 2) Mg 3) Ca 4) Sr

Topic mastery

51. Explain why hydrogen is not considered to be a member of Group 1 alkali metals.

52. Element X has an atomic radius of 160 pm, and an electronegativity of 1.3. Using the reference tables, identify the elements that X could be. Using other properties on the table, how would you test to see which of these elements you identified is element X.

53. Explain why the chemical reactivity of Group 1 elements increases from top to bottom, while it decreases from top to bottom of Group 17 elements.

54. Mendeleev arranged the Periodic Table in order of increasing atomic masses. Locate iodine and tellurium on the table and note that they are not arranged by increasing mass, and yet Mendeleev placed iodine in Group 17 and tellurium in Group 16.

    a) What is the likely reason that he did not arrange them by increasing mass?

    b) Locate two other elements on the table that are *not* arranged by increasing mass.

Lesson 1: Historical development of the modern atom

Introduction

The **atom** is the most basic unit of matter. Since atoms are very small and cannot be seen with the most sophisticated equipment, several scientists over hundreds of years have proposed different models of atoms to help explain the nature and behavior of matter.

In this lesson, you will learn about these historical scientists, their experiments and their proposed models of the atom.

Atomic models

The **wave mechanical-model** is the current and the most widely accepted model of the atom. This current model of the atom is due to work and discoveries of many scientists over hundreds of years .

According to the wave-mechanical model:

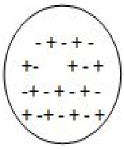
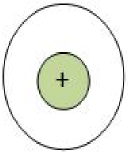
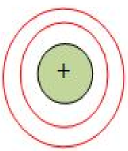
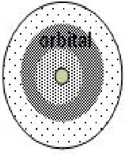
. Each atom has a small dense positive nucleus

. Electrons are found outside the nucleus in regions called **orbitals**

. An **orbital** is the most probable location of finding an electron in an atom.

Below is a list of historical scientists and their proposed models of the atom.

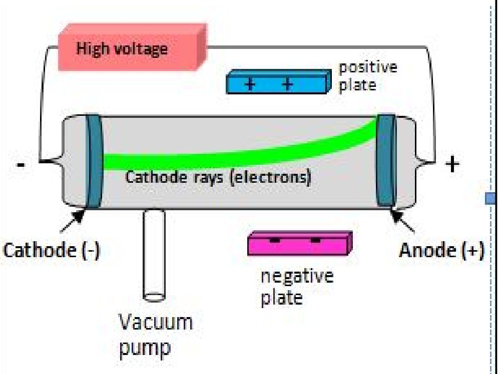
Diagrams and descriptions of each model are also given below.

John Dalton	**J.J. Thompson**	**Earnest Rutherford**	**Niels Bohr**	**Many scientists**
(earliest)				(current)
				
Hard sphere model (Cannonball model) .No internal structure	*Plum pudding model* .Electrons and positive charges disperse throughout the atom.	*Empty space model (Nuclear model)* .Small dense positive nucleus . Electrons revolve around the nucleus	*Bohr's model (Planetary model)* . Electrons in specific orbits . Orbits have fixed energy .Orbits create electron shells	*Wave mechanical (electron cloud)* .Electrons in orbitals . An orbital is the region of an atom where an electron is likely to be found

Historical Scientific Experiments

Cathode Ray experiment (JJ Thompson):

Led to the discovery of electrons



The set up

A tube with a metal disk at each end was set up to trace a beam from an electrical source. The metals were connected to an electrical source.

Anode: The Metal disk that is +.

Cathode: The Metal disk that is -

Results

A beam of light (ray) travels from the *cathode* end to the anode end of the tube. When electrically charged **+** and **-** plates were brought near the tube, the beam (ray) wss deflected toward and attracted the positive plate. The beam was repelled by the negative plate.

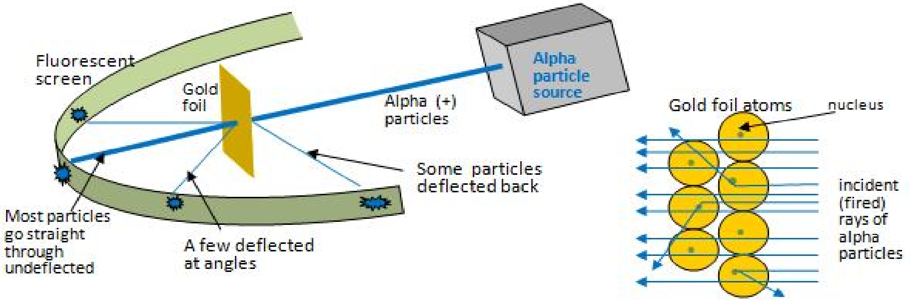
Conclusions

The beam is composed of negatively charged particles.

The term “electron” was later used to describe the negatively charged particles of an atom.

Gold Foil experiment (Rutherford)

Led to the discovery of the nucleus, and the proposed “empty space theory.”



The setup:

Alpha particles (+α) were fired at gold foil. A fluorescent screen was set up aroun the foil to detect paths of the particles once they had hit the gold foil.

Result 1

Most of the alpha particles went straight through the gold foil undeflected.

Conclusion 1

An atom is mostly empty space (Empty Space Theory)

Result 2

A few of the particles were deflected back or hit the screen at angles.

Conclusion 2

The center of the atom is dense , positive, and very small.

Lesson 2: The Atomic Structure

Introduction

Although the atom is described as the smallest unit of matter, it is also composed of much smaller particles called the ***subatomic particles***. The three ***subatomic particles*** are: proton, electron, and neutron.

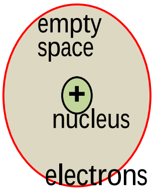
In this lesson, you will learn more about the modern atom and the subatomic particles. You will also learn the relationships between the subatomic particles, atomic number, and mass number of an atom.

Structures of atom

Atom

The atom is the basic unit of matter. All atoms (except a hydrogen atom with a mass of 1, ^{1}H) are composed of three subatomic particles: proton, electron and neutron.

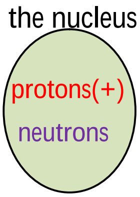
- An atom is mostly empty space
- An atom has a small dense positive core (nucleus), and negative electron cloud surrounding the nucleus
- Elements are composed of atoms with the same atomic number
- Atoms of the same element are similar
- Atoms of different elements are different



Nucleus

The nucleus is the center (core) of an atom.

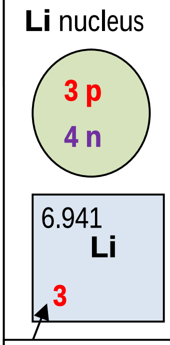
- The nucleus contains protons (+) and neutrons (no charge)
- Overall charge of the nucleus is (+) due to the protons
- Compared to the entire atom, the nucleus is small and very dense.
- Most of an atom's mass is due to the mass of its nucleus



Protons

Protons are positively charged subatomic particles found in the nucleus of an atom.

- A proton has a mass of 1 atomic mass unit (amu) and a +1 charge
- A proton is about 1836 times more massive (heavier) than an electron
- Protons are located inside the nucleus
- The number of protons is the atomic # of the element
- All atoms of the same element must have the same number of protons
- The number of protons in the nucleus is also the **nuclear charge** of the element



- Atomic #
- # of protons
- Nuclear Charge

14 Days of Question Sets for Regents and Final Exam Practice

The following section contains day-by-day practice question sets for preparing for any end-of-the-year chemistry exam.

1. Which of these terms refers to matter that could be heterogeneous?
   1) Element 2) Mixture 3) Compound 4) Solution

2. One similarity between all mixtures and compounds is that both
   1) Are heterogeneous
   2) Are homogeneous
   3) Combine in definite ratios
   4) Consist of two or more substances

3. Which correctly describes particles of a substance in the gas phase?
   1) Particles are arranged in a regular geometric pattern and are far apart
   2) Particles are in a fixed rigid position and are close together
   3) Particles move freely in a straight path
   4) Particles move freely and are close together.

4. When a substance evaporates, it is changing from
   1) Liquid to gas 2) Gas to liquid 3) Solid to gas 4) Gas to solid

5. Energy that is stored in chemical substances is called
   1) Potential energy
   2) Activation energy
   3) Kinetic energy
   4) Ionization energy

6. The specific heat capacity of water is 4.18 J/ g· ^oC. Adding 4.18 Joules of heat to a 1-gram sample of water will cause the water to
   1) Change from solid to liquid
   2) Change from liquid to solid
   3) Change its temperature 1^oC
   4) Change its temperature 4.18 ^oC

7. Real gases differ from ideal gases because the molecules of real gases have
   1) Some volume and no attraction for each other
   2) Some attraction and some attraction for each other
   3) No volume and no attraction for each other
   4) No volume and some attraction for each other

8. Under which two conditions do real gases behave most like an ideal gas?
   1) High pressure and low temperature
   2) Low pressure and high temperature
   3) High pressure and high temperature
   4) Low pressure and low temperature

9. At constant pressure, the volume of a confined gas varies
   1) Directly with the Kelvin temperature
   2) Indirectly with the Kelvin temperature
   3) Directly with the mass of the gas
   4) Indirectly with the mass of the gas

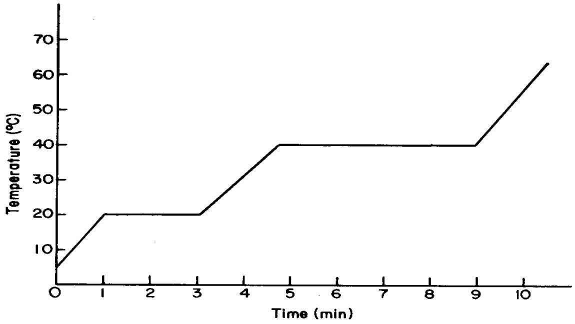
10. Under which conditions would a volume of a given sample of a gas decrease?
    1) Decrease pressure and increase temperature
    2) Decrease pressure and decrease temperature
    3) Increase pressure and decrease temperature
    4) Increase pressure and increase temperature

11. Which statement describes a chemical property of iron?
    1) Iron can be flattened into sheets.
    2) Iron conducts electricity and heat.
    3) Iron combines with oxygen to form rust.
    4) Iron can be drawn into a wire.

12. Which sample at STP has the same number of molecules as 5 liters of $NO_2(g)$ at STP?
    1) 5 grams of $H_2(g)$
    2) 5 liters of $CH_4(g)$
    3) 5 moles of $O_2(g)$
    4) 5 × 10^{23} molecules of $CO_2(g)$

13. Which substance can be decomposed by a chemical change?
    1) Ammonia 2) Potassium 3) Aluminum 4) Helium

14. The graph below represents the relationship between temperature and time as heat is added at a constant rate to a substance, starting when the substance is a solid below its melting point



During which time period does (in minutes) is the substance's average kinetic energy remain the same?
1) 0 – 1 2) 1 – 3 3) 3 - 5 4) 9 – 10

15. Molecules of which substance have the lowest average kinetic energy?
    1) NO(g) at 20°C
    2) NO_2(g) at -30°C
    3) NO_2 at 35 K
    4) N_2O_3 at 110 K

16. At STP, the difference between the boiling point and the freezing point of water in the Kelvin scale is
    1) 373 2) 273 3) 180 4) 100

17. How much heat is needed to change a 5.0 gram sample of water from 65°C to 75°C?
    1) 210 J 2) 14 J 3) 21 J 4) 43 J

18. A real gas will behave most like an ideal gas under which conditions of temperature and pressure?
    1) 0°C and 1 atm 2) 0°C and 2 atm 3) 273°C and 1 atm 4) 273°C and 2 atm

19. A 2.0 L sample of O_2(g) at STP had its volume changed to 1.5 L. If the temperature of the gas was held constant, what is the new pressure of the gas in kilopascals ?
    1) 3.0 kPa 2) 152 kPa 3) 101.3 kPa 4) 135 kPa

20. A gas occupies a volume of 6 L at 3 atm and 70°C. Which setup is correct for calculating the new volume of the gas if the temperature is changed to 150°C and the pressure is dropped to 1.0 atm?

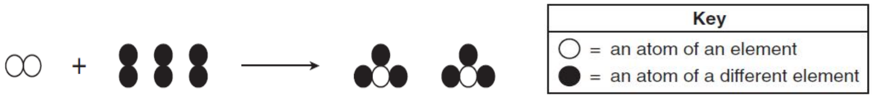
    1) $6 \times \frac{3 \times 150}{1 \times 70}$

    2) $6 \times \frac{3 \times 80}{1 \times 150}$

    3) $6 \times \frac{3 \times 423}{1 \times 343}$

    4) $6 \times \frac{3 \times 343}{1 \times 423}$

21. Given the balanced particle-diagram equation:

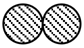


Which statement describes the type of change and the chemical properties of the product and reactants?

1) The equation represents a physical change, with the product and reactants having different chemical properties.
2) The equation represents a physical change, with the product and reactants having identical chemical properties.
3) The equation represents a chemical change, with the product and reactants having different chemical properties.
4) The equation represents a chemical change, with the product and reactants having identical chemical properties.

Constructed Responses

Bas your answers to questions 22 to 25 on the diagram of a molecule of nitrogen shown below.



represents one molecule of nitrogen.

22. Draw a particle model that shows at least six molecules of nitrogen gas.

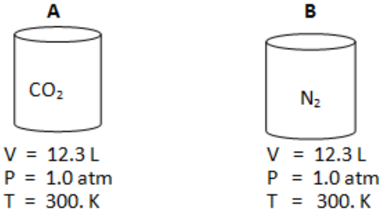
23. Draw a particle model that shows at least six molecules of liquid nitrogen.

24. Describe, in terms of particle arrangement, the difference between nitrogen gas and liquid nitrogen.

25. Good models should reflect the true nature of the concept being represented. What is the limitation of two-dimensional models ?

Base your answer to questions 26 through 28 on the information and diagrams below.

Cylinder A contains 22.0 grams of CO_2(g) and Cylinder B contains N_2(g). The volumes, pressures, and temperatures of the two gases are indicated under each cylinder.



26. How does the number of molecules of CO_2(g) in cylinder A compare to the number of molecules of N_2(g) in container B ? Your answer must include both CO_2(g) and N_2(g).

27. The temperature of CO_2(g) is increased to 450. K and the volume of cylinder A remains constant. Show a correct numerical setup for calculating the new pressure of CO_2(g) in cylinder A.

28. Calculate the new pressure of CO_2(g) in cylinder A based on your setup.

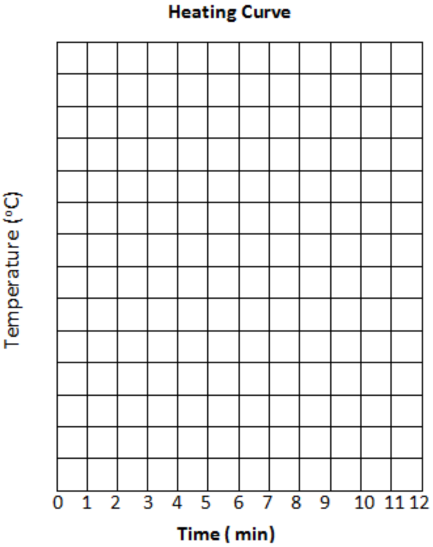
Base your answers to questions 29 through 33 on the information below.

A substance is a solid at 15°C . A student heated a sample of the substance and recorded the temperature at one-minute intervals in the data table below.

Time (min)	0	1	2	3	4	5	6	7	8	9	10	11	12
Temperature (°C)	15	32	46	53	53	53	53	53	53	53	53	60	65

29. On the grid , mark an appropriate scale on the axis labeled " Temperature (°C) ." An appropriate scale is one that allows a trend to be seen.

30 . Plot the data from the data table. Circle and connect the points.



31. Based on the data table, what is the melting point of the substance?

32. What is the evidence that the average kinetic energy of the particles of the substance is increasing during the first three minutes?

33. The heat of fusion for this substance is 122 joules per gram. How many joules of heat are needed to melt 7.50 grams of this substance at its melting point ?

1. Which determines the order of placement of the elements on the modern Periodic Table?
   1) Atomic mass
   2) Atomic number
   3) The number of neutrons, only
   4) The number of neutrons and protons

2. The elements located in the lower left corner of the Periodic Table are classified as
   1) Metals
   2) Nonmetals
   3) Metalloids
   4) Noble gases

3. The strength of an atom's attraction for the electrons in a chemical bond is measured by the
   1) density
   2) ionization energy
   3) heat of reaction
   4) electronegativity

4. What is a property of most metals?
   1) They tend to gain electrons easily when bonding.
   2) They tend to lose electrons easily when bonding.
   3) They are poor conductors of heat.
   4) They are poor conductors of electricity.

5. A metal, M, forms an oxide compound with the general formula M_2O. In which group on the Periodic Table could metal M be found?
   1) Group 1
   2) Group 2
   3) Group 16
   4) Group 17

6. Which halogen is correctly paired with its phase at STP?
   1) Br is a liquid
   2) F is a solid
   3) I is a gas
   4) Cl is a liquid

7. As the elements in Group 1 of the Periodic Table are considered in order of increasing atomic number, the atomic radius of each successive element increases. This is primarily due to an increase in the number of
   1) Neutrons in the nucleus
   2) Unpaired electrons
   3) Valence electrons
   4) Electron shells

8. When elements within Period 3 are considered in order of decreasing atomic number, ionization energy of each successive element generally
   1) Increases due to an increase in atomic size
   2) Increases due to a decrease in atomic size
   3) Decreases due to an increase in atomic size
   4) Decreases due to a decrease in atomic size

9. Which set of characteristics is true of elements in Group 2 of the Periodic Table?
   1) They all have two energy levels and have different chemical characteristics
   2) They all have two energy levels and share similar chemical characteristics
   3) They all have two valence electrons and share similar chemical properties
   4) They all have two valence electrons and have different chemical properties

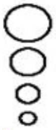
10. At STP, solid carbon can exist as graphite or as diamond. These two forms of carbon have
    1) The same properties and the same crystal structures
    2) The same properties and different crystal structures
    3) different properties and the same crystal structures
    4) different properties and different crystal structures

11. Which grouping of circles, when considered in order from the top to the bottom, best represents the relative size of the atoms of Li, Na, K, and Rb, respectively?

1)

2)

3)



4)

12. Elements strontium and beryllium both form a bond with fluorine with similar chemical formulas. The similarity in their formulas is due to
    1) Strontium and beryllium having the same number of kernel electrons
    2) Strontium and beryllium having the same number of valence electrons
    3) Strontium and beryllium having the same number of protons
    4) Strontium and beryllium having the same molecular structure

13. The element Antimony is a
    1) Metal 2) Nonmetal 3) Metalloid 4) Halogen

14. Which of these elements in Period 2 is likely to form a negative ion?
    1) Oxygen 2) Boron 3) Ne 4) Li

15. Which of these characteristics best describes the element sulfur at STP?
    1) It is brittle 2) It is malleable 3) It has luster 4) It is ductile

16. Which of these elements has the highest thermal and electrical conductivity?
    1) Iodine 2) Carbon 3) Phosphorus 4) Iron

17. Chlorine will bond with which metallic element to form a colorful compound?
    1) Aluminum 2) Sodium 3) Strontium 4)Manganese

18. According to the Periodic Table, which sequence correctly places the elements in order of increasing atomic size?
    1) Na ---> Li ----> H -----> K 3) Te ----> Sb -----> Sn ----- > In
    2) Ba ---> Sr ----> Ca-----> Mg 4) H ----> He ---> Li ----> Be

19. Which of these elements has stronger metallic characteristics than aluminum?
    1) He 2) Mg 3) Ga 4) Si

20. Which element has a greater tendency to attract electrons than phosphorus?
    1) Silicon 2) Arsenic 3) Boron 4) Sulfur

21. Which element has the greatest density at STP?
    1) barium 2) magnesium 3) beryllium 4) radium

22. An element that is malleable and a good conductor of heat and electricity could have an atomic number of
    1) 16 2) 18 3) 29 4) 35

23. Sodium atoms, potassium atoms, and cesium atoms have the same
    1) Atomic radius 3) First ionization energy
    2) Total number of protons 4) Oxidation state

24. When the elements in Group 1 are considered in order from top to bottom, each successive element at standard pressure has
    1) a higher melting point and a higher boiling point
    2) a higher melting point and a lower boiling point
    3) a lower melting point and a higher boiling point
    4) a lower melting point and a lower boiling point

25. Elements Q, X, and Z are in the same group on the Periodic Table and are listed in order of increasing atomic number. The melting point of element Q is –219°C and the melting point of element Z is –7°C. Which temperature is closest to the melting point of element X?
    1) –7°C 2) –101°C 3) –219°C 4) –226°C

Constructed Responses

Base your answer to questions 26 through 29 on the information below.

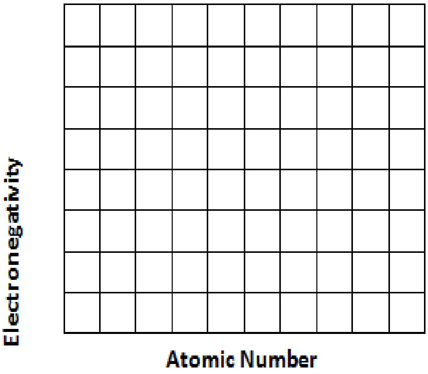
> A metal, M, was obtained from compound in a rock sample. Experiments have determined that the element is a member of Group 2 on the Periodic Table of the Elements.

26. What is the phase of element M at STP?

27. Explain, in terms of electrons, why element M is a good conductor of electricity.

28. Explain why the radius of a positive ion of element M is smaller than the radius of an atom of element M.

29. Using the element symbol M for the element, write the chemical formula for the compound that forms when element M reacts with Iodine.



Element	Atomic Number	Electronegativity
Beryllium	4	1.6
Boron	5	2.0
Carbon	6	2.6
Fluorine	9	4.0
Lithium	3	1.0
Oxygen	8	3.4

30. On the grid, set up a scale for electronegativity on the y-axis and atomic number on the x-axis. Plot the data by drawing a best-fit line.

31. Using the graph, predict the electronegativity of nitrogen.

32. For these elements, state the trend in electronegativity in terms of atomic number.

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