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Chapter 1 – Intro to Chemistry

Chemistry & Matter

Chemistry – the study of matter and its changes.

Matter & its characteristics
Matter- anything that has mass and takes up space. Mass – measurement showing the amount of matter. Weight – measure of matter and Earth’s gravitational pull. As you move away from the surface/sea level, weight is less.

What you see and what you don’t: all matter and its behavior is macroscopic – you can see it. Elements are made of particles called atoms so tiny you cannot see them – these are submicroscopic. 100 million million million atoms can fit in the period at the end of this sentence.

Branches of chemistry – because there are so many different types of matter there are many fields of chemistry.

<table>
<thead>
<tr>
<th>Branch</th>
<th>Area of emphasis</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>Organic</td>
<td>Carbon-containing chemicals</td>
<td>Plastics, medicine</td>
</tr>
<tr>
<td>Inorganic</td>
<td>Matter without carbon</td>
<td>Minerals, metals, nonmetals</td>
</tr>
<tr>
<td>Physical</td>
<td>Behavior and changes of matter and related energy changes</td>
<td>Reaction rates and mechanisms</td>
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<tr>
<td>Analytical</td>
<td>Components and composition of substances</td>
<td>Food nutrients and quality control</td>
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<td>Biochemistry</td>
<td>Matter and processes of living things</td>
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<td>Why and how chemicals interact</td>
<td></td>
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<tr>
<td>Environmental</td>
<td>Role of chemicals in the environment</td>
<td>CFCs, acid rain, PCBs</td>
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Scientific Methods

Systematic approach
Scientific method is a systematic approach used in science to do research and verify the work of others. Observations – gathering info. Qualitative data – color, shape, odor or other physical characteristics. Quantitative data – numerical info like pressure, volume, temperature, quantity of chemical formed or used up. Hypothesis – tentative explanation of observations. Experiments – set of controlled observations that test the hypothesis. Variable – a quantity or condition that can have more than one value. Independent – the variable you can change. Dependent – value changes in response to a change in the independent variable. Control – standard for comparison. Conclusion – judgment based on the info obtained. Theory – explanation supported by many experiments but still subject to new data and can be modified. Scientific law – relationship in nature supported by many experiments.
Chapter 2 – Matter: Properties and Changes

Properties of Matter

Substances – matter with uniform and unchanging composition. Ex: table salt, water.

Physical properties of matter

We can identify objects by properties – characteristics and behavior. Physical property can be observed or measured without changing the sample composition. Common properties: density, color, taste, odor, hardness, melting point, and boiling point. Extensive properties – depend on amount of substance present. Ex: mass, length and volume. Intensive properties – independent of the amount of substance. Ex: density.

Chemical properties of matter

Chemical property is the ability of a substance to combine with or change into one or more other substances. Chemical property can also be the inability to combine/change.

Observing properties of matter – every substance has its own physical and chemical properties. Observations may vary depending on the environment.


- Solid – has its own definite shape; particles are very tightly packed; expands slightly when heated.
- Liquid – flows; constant volume; takes shape of the container; less tightly packed and able to move past each other.
- Gas – flows to conform to the shape of its container and fills entire volume; particles are very far apart. Vapor is the gaseous state of substance that is solid or liquid at room temperature.

- The structure and arrangement of particles and their interactions determine the physical state of a substance at a given temperature and pressure.
Changes in Matter

Physical changes – alter a substance without changing its composition. State of matter depends on temperature and pressure of the surroundings. Ex: boil/vaporize, freeze, condense, sublime, and melt refer to phase changes.

Chemical changes – process involving one or more substances changing into new substances. Ex: explode, rust, oxidize, corrode, tarnish, ferment, burn or rot refer to chemical reactions in which the reactant substances produce different product substances.

Conservation of mass – law states that mass is neither created nor destroyed in any process, it is conserved. Mass of reactants = mass of products

A physical change results in the rearrangement of existing particles in a substance. A chemical change results in the formation of different substances with changed properties.

Elements and Compounds

Elements

All matter can be broken down into a relatively small number of building blocks called elements. Element – pure substance that cannot be separated into simpler substances by physical or chemical means. Each has a unique chemical name and symbol. 91 elements occur naturally. Dmitri Mendeleev developed the first table organizing elements according to similarities which was further developed into the periodic table. Horizontal rows = periods. Vertical columns = groups/families (have similar physical and chemical properties).

Elements are substances that are composed of atoms that have the same atomic number. Elements cannot be broken down by chemical change.

Compounds

Compound – combo of 2 or more different elements that combine chemically. Much of the matter of the universe is compounds. There are 10 million known compounds and new ones are developed and discovered at the rate of 100,000 per year. Ex: table salt is sodium chloride = sodium (Na) and chlorine (Cl) = NaCl.
Compounds can be broken down into simpler substances by chemical means. External energy (heat/electricity) is needed to separate compounds into elements. Properties of a compound differ from the elements it is made of.

- A compound is a substance composed of two or more different elements that are chemically combined in a fixed proportion. A chemical compound can be broken down by chemical means. A chemical compound can be represented by a specific chemical formula and assigned a name based on the IUPAC system.

Mixtures of Matter

Mixtures – combos of 2 or more pure substances in which each pure substance retains its individual chemical properties. Types: heterogeneous – does not blend smoothly; homogeneous – constant composition throughout and often referred to as solutions.
Mixture are composed of two or more different substances that can be separated by physical means. When different substances are mixed together, a homogeneous or heterogeneous mixture is formed.

Types of Solution Systems

<table>
<thead>
<tr>
<th>System</th>
<th>Example</th>
</tr>
</thead>
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<tr>
<td>Gas-gas</td>
<td>Air has N, O, Ar and other gases</td>
</tr>
<tr>
<td>Gas-liquid</td>
<td>Carbonated beverages is CO₂ in water</td>
</tr>
<tr>
<td>Liquid-gas</td>
<td>Moist air is water in air</td>
</tr>
<tr>
<td>Liquid-liquid</td>
<td>Vinegar is acetic acid in water</td>
</tr>
<tr>
<td>Solid-liquid</td>
<td>Powder drinks are sugar in water</td>
</tr>
<tr>
<td>Solid-solid</td>
<td>Steel is an alloy of iron with carbon</td>
</tr>
</tbody>
</table>

Separating mixtures
- Filtration – uses a porous barrier to separate a solid from a liquid. Used to separate solid from liquid.
- Distillation – technique based on differences in boiling points; can separate most homogenous mixtures. Used to separate two liquids.
- Crystallization/evaporation – results in formation of pure solid particles of a substance from a solution containing the dissolved substance. Used to separate a dissolved solid from liquid.
- Chromatography – separates components on basis of the tendency of each to travel or be drawn across the surface of another material (stationary phase). Mixture components = mobile phase. Used to separate substances into component colors.

Differences in properties such as density, particle size, molecular polarity, boiling point and freezing point, and solubility permit physical separation of the components of the mixture.

Matter is classified as a pure substance or as a mixture of substances.
Law of definite proportions – regardless of the amount, a compound is always composed of the same elements in the same proportion by mass. Percent by mass shows the ratio of mass of each element to the total mass of compound as a percentage.

- The proportions of components in a mixture can be varied. Each component in a mixture retains its original properties.
- The formula mass of a substance is the sum of the atomic masses of its atoms. The molar mass (gram formula mass) of a substance equals one mole of that substance.

### Practice – find the formula mass for each of the following

<p>| | | | |</p>
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<thead>
<tr>
<th></th>
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<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1) C₆H₁₂O₆</td>
<td>2) CCl₄</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3) H₂O</td>
<td>4) NaOH</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5) P₂O₅</td>
<td>6) CH₃COOH</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Percent by mass = (mass of element ÷ mass of compound) × 100  (Table T)

Example: compare sucrose in a sugar packet to sugar cane sucrose – shows law of definite proportions

<p>| | | |</p>
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<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Sucrose in sugar packet</td>
<td>Sugar cane sucrose</td>
</tr>
<tr>
<td></td>
<td>C</td>
<td>(8.4/20) × 100</td>
</tr>
<tr>
<td></td>
<td>H</td>
<td>(1.3/20) × 100</td>
</tr>
<tr>
<td></td>
<td>O</td>
<td>(10.3/20) × 100</td>
</tr>
<tr>
<td></td>
<td>Total</td>
<td>20.0 g</td>
</tr>
</tbody>
</table>

- A pure substance (element or compound) has a constant composition and constant properties throughout a given sample, and from sample to sample.
- The percent composition by mass of each element in a compound can be calculated mathematically.
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<td>7) 72 grams of an unknown compound contains 12 grams of hydrogen. Find percent by mass of hydrogen.</td>
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<tr>
<td>8) 80 grams of an unknown compound contains 32 grams of oxygen. Find the percent by mass of oxygen.</td>
</tr>
<tr>
<td>9) Find the percent by mass of each element in the compound C₆H₁₂O₆.</td>
</tr>
<tr>
<td>10) Find the percent by mass of each element in the compound KMnO₄.</td>
</tr>
</tbody>
</table>
Chapter 3 – Data Analysis

Units of Measurement

SI Units – standard units of the metric system. (SI = Systeme Internationale d’Unites)

Base units – based on object or event in physical world.
- Time – second (s)
- Length – meter (m)
- Mass – kilogram (kg)

Derived units – combinations of base units.
- Speed – meters/second (m/s)
- Volume – cubic meter (cm³), metric = liter (L)
- Density – compares mass to volume (g/ cm³). Density = mass / volume (Table T)

Practice – solve the following

<table>
<thead>
<tr>
<th>11) Find the density of a 13.5 g sample of aluminum that has a volume of 5 cm³.</th>
<th>12) Find the density of 22.88 g of chromium with a volume of 3.2 cm³.</th>
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<tbody>
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<td>13) What is the mass of 2 cm³ of copper?</td>
<td>14) What is the volume of 86.85 g of gold?</td>
</tr>
</tbody>
</table>

Temperature – hot and cold are qualitative, measured by temperature. There are 2 temperature scales used in chemistry: Celsius (°C) and Kelvin (K). K = °C + 273 (Table T).
- Celsius: freezing point = 0°C, boiling point = 100°C
- Kelvin: freezing point = 273 K, boiling point = 373 K

Practice – convert the following

1) Convert 32 Kelvin to Celsius.
2) Convert 95°C to Kelvin.
3) Convert 428 K to °C.
4) Convert – 20°C to Kelvin.
Scientific Notation and Dimensional Analysis

Scientific notation – expresses numbers as a multiple of 2 factors: a number between 1 and 9, and ten raised to a power (exponent)

Practice – express the following in scientific notation

1) 428 grams
2) 0.000526 cm
3) 439,000 particles
4) 0.0000000079 liters

Dimensional analysis – focuses on the units used to describe matter. Conversion factors – ratios of equivalent values used to express the same quantity in different ways.

Prefixes used with SI units

<table>
<thead>
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<th>Prefix</th>
<th>Symbol</th>
<th>Measure</th>
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<tr>
<td>Giga</td>
<td>G</td>
<td>$10^9$</td>
</tr>
<tr>
<td>Mega</td>
<td>M</td>
<td>$10^6$</td>
</tr>
<tr>
<td>Kilo</td>
<td>k</td>
<td>$10^3$</td>
</tr>
<tr>
<td>Deci</td>
<td>d</td>
<td>$10^{-1}$</td>
</tr>
<tr>
<td>Centi</td>
<td>c</td>
<td>$10^{-2}$</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Measure</th>
</tr>
</thead>
<tbody>
<tr>
<td>Milli</td>
<td>m</td>
<td>$10^{-3}$</td>
</tr>
<tr>
<td>Micro</td>
<td>μ</td>
<td>$10^{-6}$</td>
</tr>
<tr>
<td>Nano</td>
<td>n</td>
<td>$10^{-9}$</td>
</tr>
<tr>
<td>Pico</td>
<td>p</td>
<td>$10^{-12}$</td>
</tr>
</tbody>
</table>

Practice – convert the following

1) 325 g = _______________ kg
2) 0.042 g = _______________ mg
3) 425780 pm = _______________ cm
4) $9.1 \times 10^6$ mm = _______________ nm

How Reliable are Measurements?

Accuracy – how close a measured value is to an accepted value.
Precision – how close a series of measurements is to one another.

Percent error – ratio of an error to an accepted value.
Percent error = \(\frac{(\text{measured value} - \text{accepted value})}{\text{accepted value}} \times 100\)  

<table>
<thead>
<tr>
<th>Practice – solve the following</th>
</tr>
</thead>
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<tr>
<td>Copper has a gram formula mass of 63.546 g/mol and a density of 8.96 g/cm(^3).</td>
</tr>
<tr>
<td>1) A student calculated the gram formula mass of copper to be 64 g/mol. Calculate the percent error.</td>
</tr>
<tr>
<td>2) A student calculated the density of copper to be 8.75 g/cm(^3). Calculate the percent error.</td>
</tr>
</tbody>
</table>

Significant figures – scientists indicate precision of a measurement by the number of digits they report. Significant figures include all known digits plus one estimated digit.

Rules for significant figures:
1. Nonzero numbers are always significant.
2. Zeros between nonzero numbers are always significant.
3. All final zeros to the right of a decimal are significant.
4. Zero placeholders are not significant.
5. Counting numbers and defined constants have infinite number.

<table>
<thead>
<tr>
<th>Practice – how many significant figures are in each of the following?</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) 72.3</td>
</tr>
<tr>
<td>2) 60.5</td>
</tr>
<tr>
<td>3) 6.20</td>
</tr>
</tbody>
</table>
Chapter 4 – The Atom

Early Studies of Matter

- The modern model of the atom has evolved over a long period of time through the work of many scientists.

The Philosophers

Science as we know it today did not exist thousands of years ago. Scholarly thinkers, philosophers, speculated the nature of matter – many concluded matter was composed of earth, water, air and fire or that it could be divided into smaller and smaller pieces.

Greek philosopher Democritus (460-370 B.C.) was the first to propose that matter was not infinitely divisible and that it was made of atomos (tiny individual particles that could not be cut). He believed atoms could not be created, destroyed or further divided. Other philosophers criticized his ideas and Aristotle (384-322 B.C.) rejected his theory entirely. Democritus’s ideas were not science – they could not test their ideas without controlled experiments.

John Dalton (1766 – 1844)

His ideas mark the beginning of the development of modern atomic theory. His theory included:
- All matter is made of atoms.
- All atoms of an element are identical.
- Atoms cannot be created, destroyed or divided.
- Different atoms combine to form compounds.
- In chemical reactions atoms are separated, combined or rearranged.
- Atoms of different elements are different because they have different masses.

He revived and revised Democritus’s ideas after conducting scientific research.

- Atom – smallest particle of an element that retains the properties of the element.

Subatomic Particles and the Nuclear Atom

- Subatomic particles in the nucleus are protons and neutrons.
- Each atom has a nucleus, with an overall positive charge, surrounded by one or more negative electrons.

Discovering the electron

J.J. Thomson (1856 – 1940) did the cathode ray tube experiment in the late 1890s to determine the ratio of electron charge to mass. He measured the effect of both magnetic and electric fields on a cathode ray and compared that ratio to other known ratios. The mass of the charged particle was much less than hydrogen (the lightest known atom).

Thomson proposed the “plum pudding” model of the atom stating that atoms had a spherical shape with a uniformly distributed positive charge within which individually charged negative electrons were found.
Robert Millikan (1868 – 1953) determined the charge of an electron in 1909. An electron carries a charge of $-1$. Using the charge and the charge to mass ratio he discovered the mass of an electron: $9.1 \times 10^{-28}$ g.

- Electron – negatively charged particles that are part of all forms of matter.

The nuclear atom

Ernest Rutherford (1871 – 1937) conducted an experiment with a small group of other scientists to see if alpha particles (positive charge) would be deflected as they passed through thin gold foil – known as the gold foil experiment. Most passed straight through but some were deflected. This went against Thomson’s model so Rutherford created his own model. He calculated that atoms consist mostly of empty space through which electrons move and that there was a tiny, dense region in the center (the nucleus).

The concentrated positive charge in the nucleus explains the deflection of the alpha particles: the strong repulsion between the positive nucleus and the positive alpha particles caused the large deflections.

- Nucleus – tiny dense region centrally located within the atom which contains all the positive charge and virtually all of its mass.

Discovery of the protons and neutrons

By 1920, Rutherford refined the concept of the nucleus and concluded that it contains protons.

- Proton – particle carrying equal but opposite charge of an electron; positive charge of +1.

In 1932, James Chadwick (1891 – 1974) showed that the nucleus contained a neutral particle known as the neutron.

- Neutron – mass nearly equal to the proton but carries no electrical charge.

Properties of subatomic particles

<table>
<thead>
<tr>
<th>Particle</th>
<th>Symbol</th>
<th>Electrical charge</th>
<th>Relative mass</th>
<th>Actual mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electron</td>
<td>e</td>
<td>1−</td>
<td>1/1840</td>
<td>$9.11 \times 10^{-28}$</td>
</tr>
<tr>
<td>Proton</td>
<td>p</td>
<td>1+</td>
<td>1</td>
<td>$1.673 \times 10^{-24}$</td>
</tr>
<tr>
<td>Neutron</td>
<td>n</td>
<td>0</td>
<td>1</td>
<td>$1.675 \times 10^{-24}$</td>
</tr>
</tbody>
</table>

- The proton is positive and the neutron has no charge. Electrons have negative charge.

How Atoms Differ

Atomic number

Henry Moseley (1887 – 1915) discovered that atoms of each element contain a unique positive charge in their nuclei. Atomic number is the number of protons in an atom and it identifies what element the atom is and determines its position on the periodic table of elements.

Atoms are neutral so the # of protons = # of electrons (the positive and negative cancels out to make it neutral). Atomic # = # of protons = # of electrons.
Protons and electrons have equal but opposite charges. The \# of protons = \# of electrons in an atom.

Isotopes and mass number

The number of neutrons in nuclei of atoms may differ. Atoms with the same \# of protons but different \# of neutrons are called isotopes. For example there are three types of potassium. All three have 19 protons but one has 20 neutrons, one has 21 and the other has 22. In nature most elements are found as a mixture of isotopes but the relative abundance of each isotope is constant. With potassium, 93.25\% has 20 neutrons, 0.0117\% has 21 and 6.7302\% has 22.

Isotopes of the same element differ in mass (because the number of neutrons is different). More neutrons means the atom has more mass. Mass number is the sum of protons and neutrons in the nucleus of an atom. We can put the mass number after the name or symbol for an element to identify the isotope.

Isotope notation:

<table>
<thead>
<tr>
<th>Mass #</th>
<th>Element</th>
<th>Atomic #</th>
<th>Example:</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td>109</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Ag 47</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>107</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Ag 47</td>
</tr>
</tbody>
</table>

In nature, 51.84\% is silver-107 and 48.16\% is silver-109.

We can find the number of neutrons by subtracting mass \# – atomic \# = \# neutrons

<table>
<thead>
<tr>
<th>Practice – find the number of neutrons in each the following</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) Silver-107</td>
</tr>
</tbody>
</table>

Mass of individual atoms

Because the mass of subatomic particles is so small, scientists assigned the carbon-12 atom a mass of exactly 12 atomic mass units (amu). Atomic mass unit – standard to relate other atoms to; 1 amu = 1/12 the mass of a carbon-12 atom.

<table>
<thead>
<tr>
<th>Mass of subatomic particles</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Particle</strong></td>
</tr>
<tr>
<td>Electron</td>
</tr>
<tr>
<td>Proton</td>
</tr>
<tr>
<td>Neutron</td>
</tr>
</tbody>
</table>
Mass of an element does not have to be a whole number because atomic mass is the weighted average of all the naturally occurring isotopes of an element. The mass of chlorine on the periodic table is 35.453 amu. The atomic mass of chlorine-35 is 34.969 amu and its percent abundance is 75.770%. The atomic mass of chlorine-37 is 36.966 amu and its percent abundance is 24.230%. To calculate the weighted average we multiply the mass of each isotope by its percent abundance and add each mass contribution:

\[(34.969 \times 75.770\%) + (36.966 \times 24.230\%) = 26.496 + 8.957 = 35.453 \text{ amu} \]

It is sometimes easier to put given data into a table (if not already provided in table format) as follows:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Percent abundance</th>
<th>Atomic mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>chlorine-35</td>
<td>75.770%</td>
<td>34.969</td>
</tr>
<tr>
<td>chlorine-37</td>
<td>24.230%</td>
<td>36.966</td>
</tr>
</tbody>
</table>

Practice – using the data provided, calculate the atomic mass of each of the following:

1) The element lithium has two stable isotopes, lithium-6 with an atomic mass of 6.015122 and percent abundance of 7.5% and lithium-7 with an atomic mass of 7.016003 and percent abundance of 92.5%.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Percent Abundance</th>
<th>Atomic Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

2) The element Gallium has two stable isotopes, Gallium-69 and Gallium-71. Gallium-69 has a molar mass of 68.9256 g/mol and Gallium-71 has a molar mass of 70.9247 g/mol. Gallium-69 has a relative abundance of 60.11% and Gallium-71 has a relative abundance of 39.89%.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Percent Abundance</th>
<th>Atomic Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
3) Thallium has two stable isotopes, Thallium-203 and Thallium-205. Thallium-203 has a molar mass of 202.9723 g/mol and a relative abundance of 29.52%. Thallium-205 has a molar mass of 204.9744 g/mol and a relative abundance of 70.48%.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Percent Abundance</th>
<th>Atomic Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

4) Oxygen has 3 isotopes. Oxygen-16 has the amu of 15.995 and natural abundance is 99.759%. Oxygen-17 has a mass of 16.995 amu and natural abundance is 0.037%. Oxygen-18 has a mass of 17.999 amu and natural abundance is 0.204%.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Percent Abundance</th>
<th>Atomic Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

- The mass of each proton and neutron is approximately equal to 1 amu. An electron is much less massive than a proton or neutron.
- Atoms of an element with the same # of protons but different # of neutrons are called isotopes of that element.
- Average atomic mass of an element is the weighted average of the masses of its naturally occurring isotopes.
Chapter 5 – Atomic Structure: Atoms and Ions

Quantum Theory and the Atom

Bohr model of the atom

Niels Bohr (in 1913) proposed a quantum model for the hydrogen atom saying that it has only certain allowable energy states. Ground state is the lowest allowable energy state of an atom. When an atom gains energy it is in an excited state.

Bohr related energy states to the motion of the electron within the atom. The single electron in hydrogen moves around the nucleus in certain orbits and the smaller the orbit, the lower the energy state level. He assigned quantum numbers \( n \) to each orbit and calculated the radius. When an electron moves from the excited state to the ground state it loses energy in the form of spectra (light).

Excited state electron configuration must have the same total number of electrons (as shown on the Periodic Table) but one or more electrons shifts to a higher energy level. For example, the ground state electron configuration for sodium is 2-8-1. There are many excited states that can exist for each element such as 2-7-2 or 1-8-2 which add up to the total number of 11 electrons found in sodium.

### Practice – write an excited state electron configuration for each element listed

<table>
<thead>
<tr>
<th>1) Calcium</th>
<th>3) Chlorine</th>
</tr>
</thead>
<tbody>
<tr>
<td>2) Phosphorus</td>
<td>4) Oxygen</td>
</tr>
</tbody>
</table>

### Practice – determine which element and the ground state for each excited state configuration

| 1) 2-7-4 | 3) 2-6-4 |
| 2) 2-8-7-2 | 4) 1-8-5 |

<table>
<thead>
<tr>
<th>Atomic Orbit</th>
<th>Quantum #</th>
<th>Orbit Radius</th>
<th>Energy Level</th>
<th>Relative Energy</th>
</tr>
</thead>
<tbody>
<tr>
<td>1\textsuperscript{st}</td>
<td>n = 1</td>
<td>0.0529</td>
<td>1</td>
<td>( E_1 )</td>
</tr>
<tr>
<td>2\textsuperscript{nd}</td>
<td>n = 2</td>
<td>0.212</td>
<td>2</td>
<td>( E_2 = 4E_1 )</td>
</tr>
<tr>
<td>3\textsuperscript{rd}</td>
<td>n = 3</td>
<td>0.476</td>
<td>3</td>
<td>( E_3 = 9E_1 )</td>
</tr>
<tr>
<td>4\textsuperscript{th}</td>
<td>n = 4</td>
<td>0.846</td>
<td>4</td>
<td>( E_4 = 16E_1 )</td>
</tr>
<tr>
<td>5\textsuperscript{th}</td>
<td>n = 5</td>
<td>1.32</td>
<td>5</td>
<td>( E_5 = 25E_1 )</td>
</tr>
<tr>
<td>6\textsuperscript{th}</td>
<td>n = 6</td>
<td>1.90</td>
<td>6</td>
<td>( E_6 = 36E_1 )</td>
</tr>
<tr>
<td>7\textsuperscript{th}</td>
<td>n = 7</td>
<td>2.59</td>
<td>7</td>
<td>( E_7 = 49E_1 )</td>
</tr>
</tbody>
</table>
Wave mechanical model (quantum mechanical model)

Louis de Broglie (1892 – 1987) proposed the idea that accounted for fixed energy levels of Bohr’s model in 1924. Bohr’s quantized electron orbits had characteristics similar to waves. The de Broglie equation for wavelength (\(\lambda\)) is \(\lambda = \frac{h}{mv}\) (where \(m = \) mass, \(v = \) velocity, \(h = \) Planck’s constant = \(6.626 \times 10^{-34}\) J). This equation predicts that all moving particles have waves characteristics.

Heisenberg uncertainty principle

Observing an electron produces a significant unavoidable uncertainty in the position and motion of the electron. The principle states that it is impossible to know precisely both the velocity and position of a particle at the same time.

Erwin Schrödinger (1887 – 1961) derived an equation that treated hydrogen’s electron as a wave, hence the wave mechanical model. The solution to his equation is a wave function that is related to the probability of finding the electron within a particular volume around the nucleus (orbital).

An atomic orbital is the three dimensional region around the nucleus describing the most probable location of an electron. Picture the atomic orbital as a fuzzy cloud (not like planetary revolution) – the electron cloud model has a dense nucleus.

- In the wave mechanical model (electron cloud model), electrons are in orbitals defined as the regions of most probable electron location (ground state).

Hydrogen’s atomic orbitals

Bohr model assigns quantum numbers to electron orbitals. The quantum mechanical model assigns principal quantum numbers (n) that indicate relative size and energies of orbitals – as \(n\) increases the orbital becomes larger and energy level increases. So \(n\) specifies the major energy levels called principal energy levels (PEL). The lowest principal energy level is \(n = 1\) and is the ground state.

Energy sublevels are s, p, d and f according to the shape of the orbitals: s = spherical, p = dumbbell, d and f do not all have the same shape. Each orbital has at most 2 electrons. PEL 1 has 1s orbital, PEL 2 has 2s and 2p sublevels. The 2p sublevel has 3 dumbbell orbitals labeled \(2p_x\), \(2p_y\), and \(2p_z\) based on their orientation.

Hydrogen’s 4 principal energy levels

<table>
<thead>
<tr>
<th>(n)</th>
<th>Sublevel</th>
<th># of Orbitals</th>
<th>Total # of Orbitals ((n^2))</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>s</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>2</td>
<td>s</td>
<td>1</td>
<td>4</td>
</tr>
<tr>
<td></td>
<td>p</td>
<td>3</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>s</td>
<td>1</td>
<td>9</td>
</tr>
<tr>
<td></td>
<td>p</td>
<td>3</td>
<td></td>
</tr>
<tr>
<td></td>
<td>d</td>
<td>5</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>s</td>
<td>1</td>
<td>16</td>
</tr>
<tr>
<td></td>
<td>p</td>
<td>3</td>
<td></td>
</tr>
<tr>
<td></td>
<td>d</td>
<td>5</td>
<td></td>
</tr>
<tr>
<td></td>
<td>f</td>
<td>7</td>
<td></td>
</tr>
</tbody>
</table>

Maximum # of electrons in each energy level = \(2n^2\)
• Each electron in an atom has its own distinct amount of energy.
• When an electron in an atom gains a specific amount of energy, the electron is at a higher energy state (excited state).
• When an electron returns from a higher energy state to a lower one, a specific amount of energy is released. This energy may be used to identify an element.

Electron Configuration

Ground state electron configuration

Electron configuration is the arrangement of electrons. Low energy systems are more stable than high energy systems so electrons assume the arrangement giving them the lowest energy possible – the ground state electron configuration. Ground state electron configuration can be found on the Periodic Table.

Aufbau principle – each electron occupies the lowest energy orbital available.

Aufbau diagram shows energy of each sublevel. Some features:
• All orbitals related to an energy sublevel are of equal energy (look at all 3 of the 2p sublevels).
• In a multielectron atom, the energy sublevels within a PEL have different energies (the 2p are higher energy than the 2s).
• In order of increasing energy the sequence of sublevels is s, p, d and f.
• Orbitals related to sublevels within one PEL can overlap orbitals related to sublevels of another PEL (4s is lower than 3d).

Pauli exclusion principle – each electron has an associated spin (↑ is one direction and ↓ is the opposite). There is a maximum of 2 electrons per orbital but only if they have opposite spins. Paired electrons with opposite spins = ↑↓.

Hund’s rule – single electrons with the same spin must occupy equal energy orbitals before more electrons with opposite spin can occupy the same orbitals.

Orbital diagrams and electron configuration notation

2 methods can be used to represent electron configuration:
• Orbital diagram –
  \[
  \begin{array}{c}
  \square \\
  \text{unoccupied orbital}
  \end{array}
  \quad
  \begin{array}{c}
  \uparrow
  \end{array}
  \quad
  \begin{array}{c}
  \text{orbital with one electron}
  \end{array}
  \quad
  \begin{array}{c}
  \uparrow \downarrow
  \end{array}
  \quad
  \begin{array}{c}
  \text{orbital with two electrons}
  \end{array}
  \]

Example: Carbon

\[
\begin{array}{c}
\uparrow \downarrow \\
1s
\end{array}
\quad
\begin{array}{c}
\uparrow \downarrow \\
2s
\end{array}
\quad
\begin{array}{c}
\uparrow \uparrow \square
\end{array}
\quad
2p
\]

• Electron configuration notation – PEL and sublevel with superscript to indicate number of electrons in the orbital. Example: Carbon is \(1s^2\ 2s^2\ 2p^2\)
Orbital diagrams and electron configurations notation

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic #</th>
<th>Orbital Diagram</th>
<th>Electron Configuration Notation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>1</td>
<td>↑</td>
<td>1s¹</td>
</tr>
<tr>
<td>Helium</td>
<td>2</td>
<td>↑↓</td>
<td>1s²</td>
</tr>
<tr>
<td>Lithium</td>
<td>3</td>
<td>↑↓ ↑</td>
<td>1s² 2s¹</td>
</tr>
<tr>
<td>Beryllium</td>
<td>4</td>
<td>↑↓ ↑↓</td>
<td>1s² 2s²</td>
</tr>
<tr>
<td>Boron</td>
<td>5</td>
<td>↑↓ ↑↓ ↑</td>
<td>1s² 2s² 2p¹</td>
</tr>
<tr>
<td>Carbon</td>
<td>6</td>
<td>↑↓ ↑↓ ↑↓</td>
<td>1s² 2s² 2p²</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>7</td>
<td>↑↓ ↑↑ ↑↓ ↑</td>
<td>1s² 2s² 2p³</td>
</tr>
<tr>
<td>Oxygen</td>
<td>8</td>
<td>↑↓ ↑↑ ↑↑ ↑↑ ↑↓</td>
<td>1s² 2s² 2p⁴</td>
</tr>
<tr>
<td>Fluorine</td>
<td>9</td>
<td>↑↓ ↑↑ ↑↓ ↑↓ ↑</td>
<td>1s² 2s² 2p⁵</td>
</tr>
<tr>
<td>Neon</td>
<td>10</td>
<td>↑↓ ↑↑ ↑↓ ↑↑ ↑</td>
<td>1s² 2s² 2p⁶</td>
</tr>
</tbody>
</table>

Noble gas notation can be used to shorten electron configuration for higher atoms. Example: Sodium (Na) can be written in electron configuration notation as 1s² 2s² 2p⁶ 3s¹ or with noble gas notation as [Ne] 3s¹.

The sublevel diagram must be referred to in order to write the notations correctly since it follows Aufbau’s principle.

Valence electrons

Valence electrons are in the outermost orbitals (highest PEL). They determine chemical properties of the element because they are involved in bonding. You can look at the periodic table to find valence electrons for each element. Example: sulfur (S) has 6 valence, cesium (Cs) has 1 valence, francium (Fr) has 1 valence.

Electron dot structures consist of an element’s symbol (which represents the atom’s nucleus and inner electrons) and is surrounded by dots (representing the valence electrons). First two dots should be added to one of the 4 sides of the symbol (representing the s orbital), then 1 dot to each side and then a second dot to each side (the number of dots depends on the number of valence electrons – the maximum is 8).

- Outermost electrons in an atom are called valence electrons. In general, the # of valence electrons affects the chemical properties of an element.
Unstable Nuclei and Radioactive Decay

Radioactivity

Nuclear reactions involve a change in the atom’s nucleus. Atom of one element changes to another (transmutation). Radioactivity is a process in which substances spontaneously emit radiation (rays and particles) because the nuclei are unstable. Radioactive decay is a process that does not require energy; unstable nuclei lose energy by emitting radiation.

Types of radiation

Alpha radiation – radiation is deflected toward the negative. Alpha particle has 2 protons and 2 neutrons so it is a 2+ charge (equivalent to He-4) and is represented by $^4\text{He}$ or $\alpha$. Nuclear equation shows atomic # and mass # of particles involved. Ex: alpha decay of radium-226 into radon-222.

Beta radiation – deflected toward positive. Beta particles are fast moving electrons with a 1- charge and are represented by $\beta$ or $e^-$. Ex: beta decay of carbon-14 into nitrogen-14.

Gamma radiation – high energy radiation. Gamma rays have no mass and are denoted by $\gamma$. They are not deflected because they have no charge. They are usually accompanied by alpha or beta decay.

Nuclear stability depends on the ratio of neutrons to protons. Too many or too few neutrons make the atom unstable.

Practice – find the number of valence electrons and draw the electron dot structure for each

<p>| | | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1) Iron</td>
<td>3) Lead</td>
<td>5) Aluminum</td>
<td>7) Selenium</td>
</tr>
<tr>
<td>2) Phosphorus</td>
<td>4) Bromine</td>
<td>6) Calcium</td>
<td>8) Uranium*</td>
</tr>
</tbody>
</table>
Chapter 7 – The Periodic Table

Development of the Modern Periodic Table

History of the periodic table’s development
Antoine Lavoisier (1790s) compiled a list of 23 elements. By 1870 there were 70 known elements. John Newlands proposed an organization scheme for elements in 1864. He arranged elements by atomic mass and their properties repeated every 8th element so he called them octaves.
Meyer and Mendeleev demonstrated a connection between atomic mass and elemental properties. Mendeleev organized them into the first periodic tables – arranging them into columns with similar properties in order of increasing atomic mass. It was not completely correct because several elements were in groups with different properties. Moseley arranged the elements by atomic number resulting in a clear periodic pattern of properties. Periodic law – there is periodic repetition of chemical and physical properties of the elements when arranged by atomic number.

The modern periodic table
The modern periodic table consists of boxes each with an element name, symbol, atomic number and atomic mass. Columns are called groups/families and rows are called periods. Each group is numbered 1 to 18. Representative elements in groups 1, 2, 13 – 18 possess a wide range of chemical and physical properties. Transition elements are in groups 3 – 12.

Classifying the elements – 3 main classifications:
- Metals – generally shiny when smooth and clean; solid at room temperature and good conductors of heat and electricity; most are ductile (can be drawn into wire) and malleable (can be hammered into shapes).
- Nonmetals – generally gases or brittle, dull solids; poor conductors of heat and electricity (Br is the only liquid at room temperature).
- Metalloids – physical and chemical properties of both metals and nonmetals (B, Si, Ge, As, Sb, and Te).

Practice – list the atomic number and classification (metal, metalloid, nonmetal) for each

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td>Oxygen</td>
<td>3)</td>
</tr>
<tr>
<td>2)</td>
<td>Silicon</td>
<td>4)</td>
</tr>
<tr>
<td>5)</td>
<td>Cobalt</td>
<td>6)</td>
</tr>
</tbody>
</table>
The left side of the periodic table (except hydrogen) is all metals. Metalloids all touch the bolded steps (except Al, Po and At). Nonmetals are found on the right side of the periodic table. Group 1 = alkali metals, group 2 = alkaline earth metals, group 17 = halogens, and group 18 = noble gases.

- The placement or location of elements on the periodic table gives an indication of physical and chemical properties of that element. The elements on the periodic table are arranged in order of increasing atomic number.
- The number of protons in an atom (atomic number) identifies the element. The sum of protons and neutrons in an atom (mass number) identifies an isotope.
- Elements can be classified by their properties and located on the periodic table as metals, nonmetals, metalloids and noble gases.

**Classification of the Elements**

**Organizing elements by electron configuration**

<table>
<thead>
<tr>
<th>Period</th>
<th>Element</th>
<th>Electron configuration</th>
<th>Noble gas notation</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Hydrogen</td>
<td>1s^1</td>
<td>1s^1</td>
</tr>
<tr>
<td>2</td>
<td>Lithium</td>
<td>1s^2 2s^1</td>
<td>[He] 2s^1</td>
</tr>
<tr>
<td>3</td>
<td>Sodium</td>
<td>1s^2 2s^2 2p^6 3s^1</td>
<td>[Ne] 3s^1</td>
</tr>
<tr>
<td>4</td>
<td>Potassium</td>
<td>1s^2 2s^2 2p^6 3s^2 3p^6 4s^1</td>
<td>[Ar] 4s^1</td>
</tr>
</tbody>
</table>

**Valence electrons**
Atoms in the same group have similar chemical properties because they have the same number of valence electrons. The energy level of valence electrons indicates the period in which it is found.
- For groups 1, 2 and 13 – 18 on the periodic table, elements within the same group have the same number of valence electrons (He is an exception) and therefore have similar chemical properties.

**The s, p, d and f-block elements**

The periodic table has 4 distinct blocks because there are 4 sublevels. The s-block has valence electrons only in the s orbital. The p-block have partially filled p orbitals. The s and p-block comprise the representative elements. The d-block is the largest block and has transition metals with filled or partially filled d orbitals. The f-block has inner transition metals.

**Periodic Trends**

**Atomic radius**
For metals it is ½ the distance between adjacent nuclei in a crystal of the element. For nonmetals it is ½ the distance between nuclei of identical atoms chemically bonded together. Atomic radius can be found on Table S.
- In periods: atomic radius decreases as you move from left to right because there is increased charge in the nucleus.
- In groups: atomic radius increases as you move down a group from top to bottom.
Practice – record the atomic radius of each

| 1) Mg | 2) Li | 3) Sc | 4) Rb | 5) Pb |

Ionic radius
Atoms gain or lose 1 or more electrons to form ions. An ion is an atom or bonded group of atoms with a positive or negative charge. When they lose electrons they get smaller for two reasons: lost electrons are the valence electrons leaving the outer orbital empty and because electrostatic repulsion between fewer electrons decreases so they are pulled closer to the nucleus. When they gain electrons they get larger.
- In periods: from left to right the size of the + ions decreases, in 15 or 16 the larger – ion decreases.
- In groups: as you move down a group an ion’s outer electrons are in higher PELs so there is a gradual increase in ionic size.

Practice – determine if ionic radius is larger or smaller than atomic radius for each

| 1) Na | 2) Cl | 3) Ba | 4) O | 5) Mn |

Ionization energy
Ionization energy is the energy required to remove an electron from a gaseous atom. It indicates how strongly a nucleus holds onto valence electrons. High IE = strong hold, less likely to form + ions. Low IE = weak hold, more likely to form + ions. Octet rule – atoms tend to gain, lose or share electrons in order to acquire a full set of 8 valence electrons. The right side of the periodic table tends to gain electrons to form noble gas configuration. The left side tends to lose electrons forming positive ions.
- In periods: IE increases as you move left to right.
- In groups: IE decreases as you move down a group.

Practice – record the ionization energy for each

| 1) H | 2) Ca | 3) F | 4) Ne | 5) Fr |

Electronegativity
Electronegativity is the relative ability of its atoms to attract electrons in a chemical bond. Fluorine (F) is the most electronegative with EN = 3.98 and francium (Fr) is the least with EN = 0.7. Atoms with greater EN more strongly attract bond’s electrons.
- In periods: EN increases from left to right.
- In groups: EN decreases from top to bottom.
Practice – record the electronegativity for each

<p>| | | | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td>H</td>
<td>2)</td>
<td>Ca</td>
<td>3)</td>
</tr>
</tbody>
</table>

- Succession of elements within the same group demonstrates characteristic trends: differences in atomic radius, ionic radius, electronegativity, first ionization energy, metallic/nonmetallic properties.
- Succession of elements across the same period has characteristic trends as well.

**The Elements**

**Properties of s-Block Elements**

Representative elements

Groups 1, 2 and 13 – 18 are representative elements. Elements within a group have similar physical and chemical properties. Metals tend to lose electrons; the lower the IE the more reactive the metal. As atomic number increases, reactivity increases. Nonmetals gain electrons; the higher the IE the more reactive. As atomic number increases, reactivity decreases. Diagonal relationships – there are close relationships between elements in neighboring groups (Li → Mg, Be → Al, B → Si).

**Hydrogen**

It is located in group 1 but is not part of any group because it is more similar to nonmetals of group 17. It has both metallic and nonmetallic properties. There are three isotopes: protium with no neutrons (99.985%), deuterium with 1 neutron (0.015% = H-2) and tritium with 2 neutrons (radioactive = H-3).

**Group 1: alkali metals**

Group 1 metals react with water to form alkaline solutions. They easily lose a valence electron and form ions with 1+ charge. They are highly reactive and found combined with other elements in nature. They are good conductors of heat and electricity. Group 1: lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs) and francium (Fr).

**Group 2: alkaline earth metals**

Form compounds with oxygen called oxides. All except beryllium oxide form alkaline solutions when mixed with water (alkaline) and do not melt in fire (earth). They are shiny solids and are harder than group 1. They lose 2 electrons to form 2+ ions. When exposed to oxygen they form a thin oxide coating. Most compounds do not dissolve easily. Group 2: beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra).
Properties of p-Block Elements

Group 13: the boron group
They are always combined with other elements in nature. There are one metalloid and 4 metals in the group. All but thallium lose 3 electrons to form 3+ ions. Thallium forms 1+ ions (gallium and indium can form 1+ ions too). Group 13: boron (B), aluminum (Al), gallium (Ga), indium (In) and thallium (Tl).

Group 14: the carbon group
Carbon is nonmetal, silicon and germanium are metalloids, tin and lead are metals. Carbon is not representative of the others. Carbon is an allotrope – form of an element in the same physical state (solid, liquid or gas) that have different structures and properties. Ex: solid carbon is diamond, coal and graphite. Group 14: carbon (C), silicon (Si), germanium (Ge), tin (Sn), and lead (Pb).

Group 15: the nitrogen group
Nitrogen and phosphorous are nonmetals, arsenic and antimony are metalloids, and bismuth is metal. All have 5 valence electrons but they have very different chemical and physical properties. Nitrogen forms explosive compounds, phosphorous has 3 solid allotropes, antimony and bismuth expand when changing from liquid to solid. Group 15: nitrogen (N), phosphorous (P), arsenic (As), antimony (Sb) and bismuth (Bi).

Group 16: the oxygen group
Polonium is the most metallic but not typical – it is rare, radioactive and toxic. Oxygen and sulfur are nonmetals and tend to form ions with 2− charge or share two electrons to form stable configurations. Group 16: oxygen (O), sulfur (S), selenium (Se), tellurium (Te) and polonium (Po).

Group 17: the halogens
They are able to form compounds with almost all metals (making salts). Halogens differ in physical properties. Chlorine is a gas at room temperature, bromine is a liquid (evaporates easily) and iodine is a solid (but changes to vapor). Chemical behavior is similar except astatine. They are reactive nonmetals always combined with other elements. They have 7 valence electrons so they tend to share or gain 1 electron, forming an ion with a 1− charge. Group 17: fluorine (F), chlorine (Cl), bromine (Br), iodine (I) and astatine (At).

Group 18: noble gases
They are colorless and unreactive. They are among the last naturally occurring elements to be discovered. They are known for stability and have the maximum number of electrons in the outer shell (8). Group 18: helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe) and radon (Rn).

Properties of d-Block and f-Block Elements

Transition metals
Properties: electrical conductivity, luster and malleability (like other metals). Silver = best conductor of electricity; iron and titanium = great strength.
Physical properties are determined by electron configuration. Most are hard solids with relatively high melting and boiling points. Differences in properties are based on ability of unpaired electrons in d sublevel to move into valence shell. The more unpaired electrons = greater strength and higher melting and boiling points. Ex: In period 4, chromium is the hardest because it has 5 unpaired d and 1 unpaired s. From Sc to Cr increases but from Fe to Zn decreases because orbitals are filled.

Formation of ions: they can lose 2 s electrons forming ions with 2+ charge. Unpaired d can move to outer level forming ions with 3+ or higher. Transition metals form colors except white compounds with scandium, titanium or zinc.

Magnetism and metals: magnetism – able to be affected by a magnetic field. Moving electrons create magnetism but usually cancel because they have opposite spins. Diamagnetism – all electrons are paired so they are unaffected or repelled by magnetic fields. Paramagnetism – unpaired electrons in valence orbital so attracts to magnetic field.

Sources of transition metals: silver, gold, platinum and titanium = only metals unreactive enough to be found uncombined with other elements. All other transition metals are combined with nonmetals in minerals which are mixed with other materials in ores. Metallurgy – study and design ways to extract metals from compounds.

Inner transition metals

Period 6 lanthanide series: silvery metals with high melting points. They are found mixed in nature and are very hard to separate.

Period 7 actinide series: radioactive elements. Only 3 exist in nature; the rest are synthetic and are called transuranium elements (because the atomic number is greater than 92, uranium). They are created in particle accelerators or nuclear reactors. They decay quickly, except plutonium-239 (which is used as fuel in nuclear power plants).

- Elements can be differentiated by their physical properties such as density, conductivity, malleability, solubility and hardness which differ among elements.
- Elements can be differentiated by their chemical properties, which describe how elements behave during chemical reactions.
- Some elements exist in 2 or more forms in the same phase. These forms differ in molecular or crystal structure and hence their properties.
Chapter 8 – Chemical Bonding

Ionic Compounds

Chemical bonding

A chemical bond is a force that holds 2 atoms together. They may form by the attraction between a positive nucleus and negative electrons or between positive ions and negative ions. Electron dot structures are good to show bonds. IE = how easily an atom loses an electron; electron affinity = how much attraction an atom has for electrons. Noble gases have high IE and low electron affinity so there is low reactivity. Difference in reactivity is directly related to valence electrons.

Formation of positive ions occurs when an atom loses 1 or more valence electrons to attain noble gas configuration (8 valence electrons). Sodium (Na) has 1 valence electron and when it loses it, it will have the electron configuration of neon (Ne) forming a + ion; the ground state 2-8-1 becomes 2-8.

<table>
<thead>
<tr>
<th>Na atom</th>
<th>Na⁺ ion</th>
</tr>
</thead>
<tbody>
<tr>
<td>11 p⁺</td>
<td>11 p⁺</td>
</tr>
<tr>
<td>11 e⁻</td>
<td>10 e⁻</td>
</tr>
<tr>
<td>0 charge</td>
<td>+ 1 charge</td>
</tr>
</tbody>
</table>

Cation = positive ion. Na ions are more stable than Na atoms because when it loses its 1 valence it has the stable full valence shell (8 electrons). Group 1 loses 1 valence forming ions with 1+ charge. Group 2 loses 2 valence forming 2+ ions.

It is difficult to predict the number of electrons lost by transition metals. They usually form ions with 2+ or 3+ charges. Groups 11 through 14 in periods 4, 5 and 6 lose electrons to form an outer energy level with full s, p and d sublevels (pseudo noble gas configuration).

Formation of negative ions occurs when nonmetals on the right side of the periodic table have great affinity for electrons to form a stable configuration. Chlorine in group 17 has 7 valence electrons and gains 1 electron to form a –ion; the ground state 2-8-7 becomes 2-8-8.

<table>
<thead>
<tr>
<th>Cl atom</th>
<th>Cl⁻ ion</th>
</tr>
</thead>
<tbody>
<tr>
<td>17 p⁺</td>
<td>17 p⁺</td>
</tr>
<tr>
<td>17 e⁻</td>
<td>18 e⁻</td>
</tr>
<tr>
<td>0 charge</td>
<td>–1 charge</td>
</tr>
</tbody>
</table>

Anion = negative ion. When a negative ion forms –ide is added to the root name of the element. Ex: chlorine becomes a chloride ion. Phosphorous has 5 valence and can gain 3 electrons to form a phosphide ion with a 3– ion. Oxygen has 6 valence and gains 2 electrons forming oxide. Group 15 gains 3 electrons, group 16 gains 2 electrons and group 17 gains 1 electron.
When an atom gains 1 or more electrons, it becomes a negative ion and its radius increases (because it gains electrons). When an atom loses 1 or more electrons, it becomes a positive ion and radius decreases.

Atoms attain a stable valence electron configuration by bonding with other atoms. Noble gases have stable valence configuration and tend not to bond.

Electron dot diagrams (Lewis structures) can represent the valence electron arrangement in elements, compounds and ions.

Formation and Nature of Ionic Bonds

Formation of an ionic bond

An ionic bond is the electrostatic force that holds oppositely charged particles together in an ionic compound. Ionic bonds that form between metals and nonmetal oxygen form oxides; other ionic compounds are salts.

Binary compound contain only 2 different elements (metallic cation and nonmetal anion). Magnesium oxide, MgO, is binary but calcium sulfate, CaSO₄, is not because it has 3 different elements. The number of electrons lost must equal the number of electrons gained. Calcium fluoride is CaF₂ because Ca loses 2 electrons so 2 F atoms gain them.

Properties of ionic compounds

Bonds between atoms determine physical properties of the compound. Positive and negative ions are packed into regular repeating patterns forming ionic crystals. A crystal lattice is a 3-D geometric arrangement of particles. Ionic crystals vary in shape due to sizes and relative number of ions bonded. How strongly particles are bonded determines melting point, boiling point, and hardness. A large amount of energy is needed to break bonds because ionic bonds are relatively strong. Ionic crystals have high melting point and high boiling point, their color is related to structure, they are hard, rigid and brittle.

Charged particles must be free to move to conduct an electric current. Ionic solids are nonconductors but in a liquid state (dissolved in water) they conduct – they are called electrolytes.

Energy and the ionic bond: during any chemical reaction, energy is either absorbed or released. When energy is absorbed it is endothermic; when energy is released it is exothermic. Forming ionic compounds from positive and negative ions is almost always exothermic.

Lattice energy is the energy required to separate one mole of the ions of an ionic compound; it is directly related to size of the ion bonded. Smaller ions have more negative value for lattice energy (lithium is more negative than potassium). The charge of the ion also affects lattice energy – the larger the charges the more negative energy (MgO is 4 times greater than NaF).

Names and Formulas for Ionic Compounds

Formulas for Ionic Compounds

Ionic compounds contain crystals formed from many ions arranged in a pattern so formulas show the simplest ratio of ions involved called a formula unit. Ex: potassium
bromide, KBr, has a 1 to 1 ratio; magnesium chloride, MgCl₂, has a 1 to 2 ratio. Overall charge is 0 because the electrons lost equal the electrons gained.

Determining charge: binary compounds are made of a positive metal ion and negative nonmetal ion. A monatomic ion is a 1 atom ion (Mg²⁺ or Br⁻).

<table>
<thead>
<tr>
<th>Group</th>
<th>Atoms</th>
<th>Charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>H, Li, Na, K, Rb, Cs</td>
<td>1⁺</td>
</tr>
<tr>
<td>2</td>
<td>Be, Mg, Ca, Sr, Ba</td>
<td>2⁺</td>
</tr>
<tr>
<td>3</td>
<td>Sc, Y, La</td>
<td>3⁺</td>
</tr>
<tr>
<td>15</td>
<td>N, P, As</td>
<td>3⁻</td>
</tr>
<tr>
<td>16</td>
<td>O, S, Se, Te</td>
<td>2⁻</td>
</tr>
<tr>
<td>17</td>
<td>F, Cl, Br, I</td>
<td>1⁻</td>
</tr>
</tbody>
</table>

Transition metals and groups 13 and 14 can form several different positive ions. The charge of a monatomic ion is its oxidation number/oxidation state and it tells of the number of electrons transferred. Oxidation states are located in the upper right hand corner of each element square on the periodic table. Oxidation numbers determine formulas for the ionic compounds they form. In a formula for a compound, the cation is written first followed by the anion. Subscripts represent the numbers of ions of each element in the compound. The sum of oxidation numbers of a compound must be zero.

Compounds that contain polyatomic ions (more than 1 atom) act as individual ions so they do not change subscripts of atoms in the polyatomic ions – we put them in parentheses and put the subscript on the outside.

Practice – write the chemical formula for these ionic compounds

1) Potassium oxide
2) Calcium bromide
3) Aluminum sulfide
4) Sodium chromate
5) Ammonium sulfate
6) Magnesium chlorate
7) Iron (II) oxide
8) Iron (III) oxide

- When a bond is broken, energy is absorbed. When a bond is formed, energy is released.
- Physical properties of substances can be explained in terms of chemical bonds and intermolecular forces. These properties include conductivity, malleability, solubility, hardness, melting point and boiling point.

Naming ions and ionic compounds
Most polyatomic ions are oxyanions which are polyatomic ions composed of an element, usually a nonmetal, bonded to 1 or more oxygen atoms. Polyatomic ions are located on Table E of the Reference Tables. Oxyanions for nitrogen and sulfur:
• Ion with more oxygen uses the root of the nonmetal plus –ate.
• Ion with fewer oxygen uses the root of the nonmetal plus –ite.

Halogen oxyanions:
• Greatest amount of oxygen uses prefix per–, the nonmetal root, and –ate.
• One less oxygen uses the nonmetal root with –ate.
• Two less oxygen uses the nonmetal root with –ite.
• Three less oxygen uses prefix hypo– with the nonmetal root and –ite.

<table>
<thead>
<tr>
<th>Practice – name the following ions</th>
<th>1) NO $^{3-}$</th>
<th>3) SO$_{4}^{2-}$</th>
<th>5) ClO$_{4}^{−}$</th>
<th>7) ClO$_{2}^{−}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>2) NO $^{2−}$</td>
<td>4) SO$_{3}^{2−}$</td>
<td>6) ClO$_{3}^{−}$</td>
<td>8) ClO$^{−}$</td>
<td></td>
</tr>
</tbody>
</table>

General rules for naming ionic compounds:
1. Name cation first, then anion.
2. Monoatomic cations use the element name.
3. Monoatomic anions use the root of the element name plus the suffix –ide.
4. Group 1 and 2 metals have only 1 oxidation number. Transition metals and metals on the right often have more than 1 oxidation number.
5. If a compound has polyatomic ions, use the name of the ion.

<table>
<thead>
<tr>
<th>Practice – name the following ionic compounds</th>
<th>1) LiCl</th>
<th>4) MnS</th>
<th>7) CaCO$_{3}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>2) MgI$_{2}$</td>
<td>5) Mn$<em>{2}$S$</em>{7}$</td>
<td>8) Hg$<em>{2}$Cr$</em>{2}$O$_{7}$</td>
<td></td>
</tr>
<tr>
<td>3) B$<em>{2}$S$</em>{3}$</td>
<td>6) K$_{3}$N</td>
<td>9) Mg(ClO$<em>{2}$)$</em>{2}$</td>
<td></td>
</tr>
</tbody>
</table>

**Metallic Bonds and Properties of Metals**

Metallic bonds

Metals do not bond ionically but do form lattices. They do not share or lose electrons but instead a metal atom’s outer energy levels overlap. The electron sea model suggests that all metal atoms contribute to the solid’s valence electrons to form a sea of mobile electrons.
The electrons are delocalized because they are free to move. A metallic bond is the attraction of a metallic cation for delocalized electrons.

Properties

   Typical physical properties of metals can be explained by metallic bonding. They have moderately high melting point and boiling point; they are malleable and ductile; they are durable, have luster and are good conductors of heat and electricity.

Metal alloys

   An alloy is a mixture of elements that has metallic properties. The properties of alloys differ from those of the elements in the alloys. There are 2 types: substitutional and interstitial. Substitutional has atoms of original metallic solid replaced by other metal atoms of similar size. Interstitial has small holes (interstices) in metallic crystal filled with smaller atoms.

Covalent Bonding

The Covalent Bond

Why do atoms bond?

   Noble gases have stable electron configurations so they do not react with other compounds. Metals can react with nonmetals to form binary ionic compounds when electrons are transferred and the resulting ions have noble gas electron configurations. Sometimes atoms both need to gain valence electrons and have similar attraction to electrons to become stable so they share electrons. The octet rule states that atoms lose, gain or share electrons to achieve a stable configuration of 8 valence electrons (an octet).

   Chemical bonds are formed when electrons are:

      • Transferred from one atom to another (ionic).
      • Shared between atoms (covalent).
      • Mobile within a metal (metallic).

What is a covalent bond?

   A covalent bond results from the sharing of electrons. Shared electrons are part of the complete outer energy level of both atoms involved. It occurs when elements are close to each other on the periodic table, usually between nonmetals. A molecule is formed when 2 or more atoms bond covalently.

   Formation of covalent bonds: hydrogen (H₂), nitrogen (N₂), oxygen (O₂), fluorine (F₂), chlorine (Cl₂), bromine (Br₂), and iodine (I₂) occur as diatomic molecules in nature because they are more stable. The most stable arrangement of atoms exists at the point of maximum attraction. Each fluorine has 1 bonding pair of electrons and 3 lone pairs.

Single covalent bonds

   When a single pair of electrons is shared a single covalent bond forms. Shared electron pair (the bonding pair) is represented either by a pair of dots or a line in Lewis structures. Ex: a hydrogen molecule would be H:H or H–H and its molecular formula is H₂.
Halogens in group 17 have 7 valence electrons so they share 1 pair of electrons to
form an octet in a single covalent bond. Group 16 shares 2 electrons to form 2 covalent
bonds. Group 15 forms 3 covalent bonds and group 14 forms 4 covalent bonds.
The sigma bond (σ) = single covalent bond that occurs when electrons are shared in
an area centered between 2 atoms. Sigma bonds result when orbitals overlap end to end.

Multiple covalent bonds
Many molecules attain noble gas configuration by sharing more than 1 pair of
electrons between 2 atoms forming a multiple covalent bond such as a double bond or triple
The pi bond (π) = parallel orbitals overlap to share electrons. A double covalent bond
has 1 sigma and 1 pi bond. A triple covalent bond has 1 sigma and 2 pi bonds. Pi bonds are
always with sigma bonds in multiple covalent bonds.

Strength of covalent bonds
Covalent bonds involve repulsive and attractive forces; in a molecule, nuclei and
electrons attract each other but the nucleus of one molecule repels other nuclei and electrons
repel other electrons. Bonds can be broken when this balance is upset.
Strength of the bond depends on how much distance separates bonded nuclei. Bond
length is the distance between 2 bonding nuclei at the position of maximum attraction and is
determined by size of atoms and how many electrons are shared. Bond dissociation energy is
the amount of energy needed to break a specific covalent bond; it indicates strength.
In chemical reactions, bonds in reactants are broken and new bonds form as products.
Total energy change is determined by energy of bonds broken and formed. Endothermic =
greater amount of energy is needed to break existing bonds in reactants than is released.
Exothermic = more energy is released forming new bonds that is required to break bonds in
the reactants.

Naming Molecules

Naming binary molecular compounds
Binary molecular compounds are composed of 2 different nonmetals (no metals or
ions). Many have common names as well as scientific names. Rules:
1. The first element in the formula is named first using the entire element name.
2. The second element is named using the root plus the suffix –ide.
3. Prefixes indicate the number of atoms of each type element in the compound.
4. Remember that mono– is never used for the first element.

<table>
<thead>
<tr>
<th># of atoms</th>
<th>Prefix</th>
<th># of atoms</th>
<th>Prefix</th>
<th># of atoms</th>
<th>Prefix</th>
<th># of atoms</th>
<th>Prefix</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Mono</td>
<td>4</td>
<td>Tetra</td>
<td>6</td>
<td>Hexa</td>
<td>9</td>
<td>Nona</td>
</tr>
<tr>
<td>2</td>
<td>Di</td>
<td>5</td>
<td>Penta</td>
<td>7</td>
<td>Hepta</td>
<td>10</td>
<td>Deca</td>
</tr>
<tr>
<td>3</td>
<td>Tri</td>
<td></td>
<td></td>
<td>8</td>
<td>Octa</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Practice – name the following molecular compounds

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td>H₂O</td>
<td>4)</td>
</tr>
<tr>
<td>2)</td>
<td>CO</td>
<td>5)</td>
</tr>
<tr>
<td>3)</td>
<td>CCl₄</td>
<td>6)</td>
</tr>
<tr>
<td>7)</td>
<td>C₃N₂</td>
<td>8)</td>
</tr>
<tr>
<td>9)</td>
<td>B₁₃</td>
<td></td>
</tr>
</tbody>
</table>

Naming Acids
Water solutions of some molecules are acidic and are named as acids. If the compound produces hydrogen ions (H⁺) in solution it is considered an acid. There are 2 types of acids: binary acids and oxyacids.

Naming binary acids: they contain hydrogen and one other element. We use the prefix hydro– to name the hydrogen part and the rest of the name has the root of the second element plus –ic acid. As long as no oxygen is present it is considered binary and named that way. Ex: HCN = hydrocyanic acid; H₂S = hydrosulfuric acid.

Naming oxyacids: they contain oxyanion (an ion with oxygen) and hydrogen. First identify the anion and use its root with the correct suffix and acid. When the anion ends with –ate we use –ic acid; when it ends with –ite we use –ous acid. Ex: HNO₃ = nitric acid; HNO₂ = nitrous acid; HClO₃ = chloric acid; HClO₂ = clorouric acid.

- In a multiple covalent bond, more than 1 pair of electrons is shared between the 2 atoms.
- Metals tend to react with nonmetals to form ionic compounds. Nonmetals tend to react with other nonmetals to form molecular (covalent) compounds. Ionic compounds with polyatomic ions have both ionic and covalent bonding (ionic between the polyatomic ion and element; covalent between the elements in the polyatomic ion).

Practice – name the following acids

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td>HCl</td>
<td>3)</td>
</tr>
<tr>
<td>2)</td>
<td>H₃PO₄</td>
<td>4)</td>
</tr>
<tr>
<td>5)</td>
<td>H₂CO₃</td>
<td>6)</td>
</tr>
</tbody>
</table>
Molecular Structure

Structural formulas

Use letter symbols and bonds to show the relative positions of atoms. Lewis structures help predict structural formulas. Steps:

1. Predict the location of certain atoms: a) hydrogen is always a terminal (end) atom; b) the atom with the least attraction for shared electrons in the molecule is the central atom (usually the left side of the periodic table – lower EN).
2. Find the total # of electrons available for bonding.
3. Determine the # of bonding pairs by dividing # of electrons available for bonding by 2.
4. Place 1 bonding pair between the central atom and each terminal atom.
5. Subtract pairs in step 4 from pairs in step 3. The remaining electrons include lone pairs and those for double and triple bonds. Lone pairs go to terminal atoms and the rest go to central atoms.
6. If the central atom does not have 4 pairs it is not an octet. Convert 1 or 2 lone pairs on the terminal atoms to form double or triple bonds.
7. For polyatomic ions there are more or less electrons because of the positive or negative charge. Find the # of electrons and subtract the ion charge.

Resonance structures

More than one valid Lewis structure can be written for certain compounds/ions. Common = O₃, NO₃⁻, NO₂⁻, SO₃²⁻ and CO₃²⁻.

Exceptions to the octet rule

A small group of compounds have odd # of valence electrons and cannot form an octet. Ex: NO₂, ClO₂ and NO. Some compounds form with fewer than 8 valence electrons present but this is rare. Ex: BH₃. Coordinate covalent bonds form when 1 atom donates a pair of electrons to be shared with an atom or ion that needs 2. Elements with more than 8 valence electrons have an expanded octet. Ex: PCl₅, SF₆ and XeF₄.

Molecular Shape

VSEPR model

The Valence Shell Electron Pair Repulsion (VSEPR) model determines molecular shape. It is based on the arrangement that minimizes repulsion of shared and unshared electrons around a central atom. The bond angle is the angle formed between 2 terminal and 1 central atom. Lone pairs have larger orbitals because they are not shared.

Hybridization

Hybrids form from combining 2 of the same type of objects and they have characteristics of both. Hybridization – atomic orbitals are mixed to form new identical hybrid orbitals.

Electronegativity and Polarity
Electronegativity difference and bond character

Electron affinity increases as atomic number increases across periods but decreases down a group. EN difference can predict character and type of bond. EN difference is 0 when nonpolar covalent because electrons are equally shared. Unequal sharing results in polar covalent bonds. As difference in EN increases, the bond is more ionic. Ionic bonds occur when EN difference is 1.70 or greater; covalent occurs when EN difference is less than 1.70.

Polar covalent bonds

A polar bond (dipole) has more EN at the partially negative end and less EN at the partially positive end. Molecular polarity – molecules are either polar or nonpolar depending on location and nature of covalent bonds. Nonpolar molecules are not attracted to an electric field. Because charge is equally distributed, nonpolar molecules have symmetry. Polar molecules align with an electric field because there is greater electron density on one side. Because charge is not evenly distributed, polar molecules do not have symmetry (asymmetrical).

Solubility: polar molecules and ionic compounds are usually soluble in polar substances. Nonpolar dissolve only in nonpolar substances. Remember: like dissolves like.

Properties of covalent compounds

Differences in properties result in difference in attractive forces. In covalent compounds the bond between atoms is strong but the attractions between molecules (intermolecular/Van der Waals forces) are weak.

For nonpolar molecules the attraction between molecules is called a dispersion force or induced dipole and is weak. For polar molecules the attraction is called dipole-dipole force and is stronger; it occurs between one end of a dipole and the oppositely charged end of another.

Physical properties are due to intermolecular forces. Melting point and boiling point are low compared to ionic substances. Many exist as gases or vaporize at room temperature. They are relatively soft solids because the forces are weak.

Covalent network solids

Network solids are composed only of atoms interconnected by a network of bonds. Ex: quartz and diamond. They are typically brittle, nonconductors and extremely hard.

- Compounds can be differentiated by chemical and physical properties.
- Two major categories of compounds are ionic and molecular (covalent).
- Molecular polarity can be determined by the shape of the molecule and the distribution of charge. Symmetrical (nonpolar) molecules include CO\textsubscript{2}, CH\textsubscript{4} and diatomic elements. Asymmetrical (polar) molecules include HCl, NH\textsubscript{3}, and H\textsubscript{2}O.
- EN indicates how strongly an atom of an element attracts electrons in a chemical bond. EN values are assigned according to arbitrary scales.
- EN difference between 2 bonded atoms is used to assess the degree of polarity in the bond.
Chapter 10 – Chemical Reactions

Reactions and Equations

Evidence of chemical reactions
A chemical reaction is a process by which atoms of one or more substances are rearranged to form different substances. Evidence: temperature change (release or absorb energy), color change, odor, gas release/bubbles, appearance of a solid.

Representing chemical reactions
Equations represent chemical reactions and show reactant (the starting substances) and products (substances formed during reactions). Equations show direction of the reaction with an arrow → or ↔. The arrow separates reactants (R) from products (P) and when reading an equation the arrow means “yields”. Ex: reactant 1 + reactant 2 → reactant 3 + reactant 4.

It is important to show the physical state of R and P. Symbols used in equations: (s) means solid, (l) means liquid, (g) means gas and (aq) means aqueous which is a substance dissolved in water.

Word equations indicate R and P using the names of the substances involved (not the chemical symbols). Ex: R1 + R2 → P1 could be shown as iron (s) + chlorine (g) → iron (III) chloride (s).

Skeleton equations use chemical formulas for R and P (instead of words) but are not balanced properly. Ex: Fe (s) + Cl₂ (g) → FeCl₃ (s).

Chemical equations obey the law of conservation of mass because the R and P are properly balanced. Ex: 2 Fe (s) + 3Cl₂ (g) → 2FeCl₃ (s).

❖ In all chemical reactions there is a conservation of mass, energy, and charge.
❖ A balanced chemical equation represents conservation of atoms. The coefficients in a balanced chemical equation can be used to determine mole ratios in the reaction.

Balancing chemical equations
Balanced equations reflect the law of conservation of mass. You must find the correct coefficients for the chemical formulas in the skeleton equation. Coefficient = the # in front of a chemical formula, usually whole numbers, tells the smallest # of particles of a substance in a reaction. Coefficients describe the lowest whole # ratio of R to P.

Steps for balancing:
1. Write the skeleton equation.
2. Count the atoms of elements in the reactants.
3. Count the atoms of elements in the products.
4. Change coefficients to make the # of atoms of each element equal on both sides of the equation.
5. Write the coefficients in the lowest possible ratio.
6. Check your work.
Practice – balance the following equations

1) \( \_\_\_\text{Al} + \_\_\_\text{O}_2 \rightarrow \_\_\_\text{Al}_2\text{O}_3 \)

2) \( \_\_\_\text{Fe}_2\text{O}_3 + \_\_\_\text{Al} \rightarrow \_\_\_\text{Al}_2\text{O}_3 + \_\_\_\text{Fe} \)

3) \( \_\_\_\text{Mg} + \_\_\_\text{HCl} \rightarrow \_\_\_\text{MgCl}_2 + \_\_\_\text{H}_2 \)

4) \( \_\_\_\text{Fe} + \_\_\_\text{V}_2\text{O}_3 \rightarrow \_\_\_\text{Fe}_2\text{O}_3 + \_\_\_\text{VO} \)

5) \( \_\_\_\text{Na} + \_\_\_\text{H}_2\text{O} \rightarrow \_\_\_\text{NaOH} + \_\_\_\text{H}_2 \)

6) \( \_\_\_\text{C}_6\text{H}_{12}\text{O}_6 + \_\_\_\text{O}_2 + \_\_\_\text{H}_2\text{O} \rightarrow \_\_\_\text{CO}_2 + \_\_\_\text{H}_2\text{O} \)
Classifying Chemical Reactions

5 basic types: synthesis, combustion, decomposition, single replacement and double replacement.

Synthesis reactions

Two or more substances react to produce a single product. The general equation for synthesis would be \( A + B \rightarrow AB \). Ex: \( 2\text{Fe} (s) + 3\text{Cl}_2 (g) \rightarrow 2\text{FeCl}_3 (s) \). Synthesis occurs when 2 elements react to make a compound, when 2 compounds combine to form a larger compound, or when an element and a compound react to form a compound.

Combustion reactions

Combustion occurs when oxygen combines with a substance and releases energy in the form of heat and light. It can occur between hydrogen and oxygen when hydrogen is heated, water is formed and an explosive amount of energy is released. Ex: \( 2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \). It can be between carbon and oxygen when coal is burned. Ex: \( \text{C} + \text{O}_2 \rightarrow \text{CO}_2 \). Not all combustion reactions are synthesis; some can be replacement such as between methane and oxygen. All hydrocarbons (made only of hydrogen and carbon) burn in oxygen to yield carbon dioxide and water.

Decomposition reactions

Decomposition is the opposite of synthesis. A single compound breaks down into 2 or more elements or new compounds. The general equation is \( AB \rightarrow A + B \). Ex: \( 2\text{NaN}_3 (s) \rightarrow 2\text{Na} (s) + 3\text{N}_2 (g) \). Decomposition reactions often require an energy source such as heat, light or electricity.

Replacement reactions

Replacement reactions involve replacement of an element in a compound. There are 2 types: single and double replacement reactions.

Single replacement occurs when atoms of one element replace atoms of another element in a compound. The general equation is \( A + BX \rightarrow AX + B \). Ex: \( 2\text{Li} (s) + 2\text{H}_2\text{O} (l) \rightarrow 2\text{LiOH} (aq) + \text{H}_2 (g) \).

Metal’s reactivity is its ability to react with another substance. You can predict reactions by referring to Table J Activity Series. A metal can replace any metal listed below it that is in a compound but cannot replace any metals listed above it. We use NR to indicate that a reaction will not occur based on reactivity (NR = no reaction). Ex: \( \text{Ag} (s) + \text{Cu(NO}_3\text{)}_2 (aq) \rightarrow \text{NR} \).

Double replacement occurs when there is an exchange of ions between 2 ionic compounds (exchange of positive ions between 2 compounds in water). The positive and negative ions of 2 compounds switch places. The general equation is \( AX + BY \rightarrow AY + BX \). Ex: \( \text{Ca(OH)}_2 (aq) + 2\text{HCl} (aq) \rightarrow \text{CaCl}_2 (aq) + 2\text{H}_2\text{O} (l) \). The anions \( \text{OH}^- \) and \( \text{Cl}^- \) switched and bonded to other cations \( \text{Ca}^{2+} \) and \( \text{H}^+ \).

A solid produced during a reaction in a solution is called a precipitate. All double replacement reactions produce precipitate, gas or water as a product.
Types of chemical reactions include synthesis, decomposition, single replacement, and double replacement.

<table>
<thead>
<tr>
<th>Practice – classify the following reactions</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) $2\text{NaOH} + \text{CuCl}_2 \rightarrow 2\text{NaCl} + \text{Cu(OH)}_2$</td>
</tr>
<tr>
<td>2) $\text{H}_2\text{CO}_3 \rightarrow \text{H}_2\text{O} + \text{CO}_2$</td>
</tr>
<tr>
<td>3) $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$</td>
</tr>
<tr>
<td>4) $\text{Cu} + 2\text{AgNO}_3 \rightarrow 2\text{Ag} + \text{Cu(NO}_3)_2$</td>
</tr>
<tr>
<td>5) $2\text{SO}_2 + \text{O}_2 \rightarrow 2\text{SO}_3$</td>
</tr>
<tr>
<td>6) $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$</td>
</tr>
<tr>
<td>7) $\text{NH}_4\text{NO}_3 \rightarrow \text{N}_2\text{O} + 2\text{H}_2\text{O}$</td>
</tr>
<tr>
<td>8) $\text{F}_2 + 2\text{NaBr} \rightarrow 2\text{NaF} + \text{Br}_2$</td>
</tr>
<tr>
<td>9) $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$</td>
</tr>
<tr>
<td>10) $\text{HCl} + \text{NaHCO}_3 \rightarrow \text{H}_2\text{CO}_3 + \text{NaCl}$</td>
</tr>
</tbody>
</table>

Reactions in Aqueous Solutions

Aqueous solutions

An aqueous solution is one in which the solvent is water. There are many probable solutes that can form molecules or ions in solution. Ex: $\text{HCl} (\text{g}) \rightarrow \text{H}^+ (\text{aq}) + \text{Cl}^- (\text{aq})$. Ex: $\text{NaOH} (\text{aq}) \rightarrow \text{Na}^+ (\text{aq}) + \text{OH}^- (\text{aq})$.

Reactions that form precipitates

Chemists use ionic equations to show details of reactions that involve ions in aqueous solutions. A complete ionic equation shows all particles in a solution as they realistically exist. Spectator ions do not participate in a reaction. Net ionic equations show only the particles that participate in the reaction. You write net ionic equations by writing complete ionic equations and then crossing out all spectator ions.

Chemical equation: $2\text{NaOH} (\text{aq}) + \text{CuCl}_2 (\text{aq}) \rightarrow 2\text{NaCl} (\text{aq}) + \text{Cu(OH)}_2 (\text{s})$.

Ionic equation: $2\text{Na}^+ (\text{aq}) + 2\text{OH}^- (\text{aq}) + \text{Cu}^{2+} (\text{aq}) + 2\text{Cl}^- (\text{aq}) \rightarrow 2\text{Na}^+ (\text{aq}) + 2\text{Cl}^- (\text{aq}) + \text{Cu(OH)}_2 (\text{s})$.

Remove spectator ions: $2\text{Na}^+ (\text{aq}) + 2\text{OH}^- (\text{aq}) + \text{Cu}^{2+} (\text{aq}) + 2\text{Cl}^- (\text{aq}) \rightarrow 2\text{Na}^+ (\text{aq}) + 2\text{Cl}^- (\text{aq}) + \text{Cu(OH)}_2 (\text{s})$.

Net ionic equation: $2\text{OH}^- (\text{aq}) + \text{Cu}^{2+} (\text{aq}) \rightarrow \text{Cu(OH)}_2 (\text{s})$. 

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Reactions that form water

Water molecules produced in a reaction increase the # of solvent particles. No evidence of chemical reactions is observable because water is colorless and odorless and makes up most of the solution.

Chemical equation: \( \text{HBr (aq)} + \text{NaOH (aq)} \rightarrow \text{NaBr (aq)} + \text{H}_2\text{O (l)}. \)
Ionic equation: \( \text{H}^+ (\text{aq}) + \text{Br}^- (\text{aq}) + \text{Na}^+ (\text{aq}) + \text{OH}^- (\text{aq}) \rightarrow \text{Na}^+ (\text{aq}) + \text{Br}^- (\text{aq}) + \text{H}_2\text{O (l)}. \)
Net ionic equation: \( \text{H}^+ (\text{aq}) + \text{OH}^- (\text{aq}) \rightarrow \text{H}_2\text{O (l)}. \)

Reactions that form gases

Common gases produced: carbon dioxide, hydrogen cyanide and hydrogen sulfide.

Chemical equation: \( 2\text{HI (aq)} + \text{Li}_2\text{S (aq)} \rightarrow 2\text{LiI (aq)} + \text{H}_2\text{S (g)}. \)
Ionic equation: \( 2\text{H}^+ (\text{aq}) + 2\text{I}^- (\text{aq}) + 2\text{Li}^+ (\text{aq}) + \text{S}^{2-} (\text{aq}) \rightarrow 2\text{Li}^+ (\text{aq}) + 2\text{I}^- (\text{aq}) + \text{H}_2\text{S (g)}. \)
Net ionic equation: \( 2\text{H}^+ (\text{aq}) + \text{S}^{2-} (\text{aq}) \rightarrow \text{H}_2\text{S (g)}. \)

When you combine any acidic solution with sodium hydrogen carbonate 2 reactions occur almost simultaneously – double replacement and decomposition (similar to vinegar reacting with baking soda). The two reactions can be combined and represented by 1 equation.

Chemical equation: \( \text{HCl (aq)} + \text{NaHCO}_3 (\text{aq}) + \text{H}_2\text{CO}_3 (\text{aq}) \rightarrow \text{H}_2\text{CO}_3 (\text{aq}) + \text{NaCl (aq)} + \text{H}_2\text{O (l)} + \text{CO}_2 (\text{g}). \)
Overall: \( \text{HCl (aq)} + \text{NaHCO}_3 (\text{aq}) \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)} + \text{CO}_2 (\text{g}). \)
Ionic equation: \( \text{H}^+ (\text{aq}) + \text{Cl}^- (\text{aq}) + \text{Na}^+ (\text{aq}) + \text{HCO}_3^- (\text{aq}) \rightarrow \text{Na}^+ (\text{aq}) + \text{Cl}^- (\text{aq}) + \text{H}_2\text{O (l)} + \text{CO}_2 (\text{g}). \)
Net ionic equation: \( \text{H}^+ (\text{aq}) + \text{HCO}_3^- (\text{aq}) \rightarrow \text{H}_2\text{O (l)} + \text{CO}_2 (\text{g}). \)
Chapter 9A – The Mole

Measuring Matter

Counting particles
Chemists created their own counting unit called the mole because it is impossible to actually count the # of atoms or molecules. The mole (abbreviated mol) is the SI base unit used to measure the amount of substance. It is the # of representative particles (carbon atoms) in exactly 12 grams of pure carbon-12. A mole of anything contains \( \frac{6.0221367 \times 10^{23}}{} \) representative particles (atoms, molecules, formula units, electrons or ions) which is called Avogadro’s number, named after Amedeo Avogadro who in 1811 determined the volume of one mole of gas. We round it to \( 6.02 \times 10^{23} \).

Converting moles to particles and particles to moles:
- # of representative particles = # moles × \( 6.02 \times 10^{23} \) representative particles.
- # of moles = # representative particles ÷ \( 6.02 \times 10^{23} \) representative particles.

<table>
<thead>
<tr>
<th>Practice – use Avogadro’s number to find the following</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Moles to particles</strong></td>
</tr>
<tr>
<td>1) Atoms in 2.5 mol Zn</td>
</tr>
<tr>
<td>2) Formula units in 3.25 mol AgNO₃</td>
</tr>
<tr>
<td>3) Molecules in 12.5 mol H₂O</td>
</tr>
</tbody>
</table>

Mass and the Mole

The mass of a mole
Isotope carbon-12 is used as the standard for atomic mass. Atomic mass on the periodic table is not an integer because it is the weighted average of masses of all naturally occurring isotopes of that element.

Mass of atoms are established relative to the mass of carbon-12. Molar mass is the mass in grams of one mole of any pure substance. Molar mass is numerically equal to atomic mass and has the units g/mol. Ex: an atom of manganese has an atomic mass of 54.94 amu
so the molar mass is 54.94 g/mol. If you have 54.94 g of manganese on a balance you indirectly count 6.02 × 10^{23} atoms of manganese.

Using molar mass
- # of moles × # grams = mass

Ex: If you need 3 moles of Mn, you multiply by the mass of Mn on the Periodic Table so 3 × 54.94 g/mol = 165 grams.

<table>
<thead>
<tr>
<th>Practice – determine the mass or moles of each</th>
<th>Mole to mass</th>
<th>Mass to mole</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) 3.57 moles of Al</td>
<td></td>
<td>4) 25.5 grams of Ag</td>
</tr>
<tr>
<td>2) 42.6 moles of Si</td>
<td></td>
<td>5) 300 grams of S</td>
</tr>
<tr>
<td>3) 2.45 moles of Zn</td>
<td></td>
<td>6) 1.00 kilograms of Fe</td>
</tr>
</tbody>
</table>

Mass to atoms and atoms to mass
When changing from mass to atoms and atoms to mass, you must first find the number of moles since there is no equation to convert directly from one to the other.
- Given mass × (1 mole ÷ # grams) = # of moles THEN moles × 6.02 × 10^{23} atoms = # of atoms
- Given # of atoms × (1 mole ÷ 6.02 × 10^{23}) = # of moles THEN moles × # grams = mass
Practice – solve the following; remember you must convert to moles before mass to atoms or atoms to mass calculations

1) How many atoms are in 55.2 grams of Li?

2) How many atoms are in 0.230 grams of Pb?

3) What is the mass of \(6.02 \times 10^{24}\) atoms of Bi?

4) What is the mass of \(3.40 \times 10^{22}\) atoms of He?

Moles of Compounds

Chemical formulas and the mole

Chemical formulas indicate the types of atoms and # of each contained in one unit of the compound. Ex: Freon is \(\text{CCl}_2\text{F}_2\) which has 1 carbon, 2 chlorine and 2 fluorine. You may need to convert from moles of a compound to moles of individual atoms or from moles of individual atoms to moles of a compound depending on the question.

Ex: How many moles of F atoms are in 5.5 moles of Freon \((\text{CCl}_2\text{F}_2)\)? There are 2 mol F in 1 mol \(\text{CCl}_2\text{F}_2\), so 5.5 mol \(\times 2 = 11\) mol F atoms.

Practice – find the number of moles of atoms in the following compound

1) Moles of Cl in 2.5 mol \(\text{ZnCl}_2\)  
2) Moles of Zn in 2.5 mol \(\text{ZnCl}_2\)

Molar mass of compounds
The mass of a mole of the compound equals the sum of the masses of every particle in the compound. To find molar mass of compounds, add the masses of all elements in it.

- # of moles of atoms × molar mass = # of grams.

Ex: potassium chromate, K₂CrO₄.

\[
2 \text{ mol K} \times 39 \text{ g/mol} = 78 \text{ grams} \\
1 \text{ mol Cr} \times 52 \text{ g/mol} = 52 \text{ grams} \\
4 \text{ mol O} \times 16 \text{ g/mol} = 64 \text{ grams} \\
\text{Total molar mass} = 194 \text{ grams}
\]

<table>
<thead>
<tr>
<th>Practice – calculate the molar mass of the following compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) AlCl₃</td>
</tr>
<tr>
<td>2) (C₃H₅)₂S</td>
</tr>
<tr>
<td>3) C₆H₁₂O₆</td>
</tr>
<tr>
<td>4) Ca(OH)₂</td>
</tr>
</tbody>
</table>

Converting between moles of a compound and mass

Mole to mass conversion: Calculate molar mass of the compound then convert moles to grams by multiplying by the number of moles.

Mass to mole conversion: Calculate molar mass of the compound; take the given mass and divide by molar mass to find moles.

<table>
<thead>
<tr>
<th>Practice – convert the following</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mole to mass</td>
</tr>
<tr>
<td>----------------</td>
</tr>
<tr>
<td>1) Find the mass of 0.4 mol AlCl₃</td>
</tr>
<tr>
<td>2) Find the mass of 2.5 mol (C₃H₅)₂S</td>
</tr>
</tbody>
</table>
Empirical Formulas and Molecular Formulas

- Types of chemical formulas include empirical, molecular, and structural.

Percent composition
Percent by mass of each element in a compound is percent composition.
- \( \frac{\text{mass of element}}{\text{mass of compound}} \times 100 = \% \text{ by mass} \)

Ex: Molar mass of \( \text{H}_2\text{O} \) is 18 g/mol.
- \% by mass of H = \( \frac{2}{18} \times 100 = 11.1\% \)
- \% by mass of O = \( \frac{16}{18} \times 100 = 88.89\% \)

Practice – find the percent by mass of each element in the following compounds

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td>AlCl(_3)</td>
</tr>
<tr>
<td>2)</td>
<td>((\text{C}_3\text{H}_5)\text{S})</td>
</tr>
<tr>
<td>3)</td>
<td>(\text{C}<em>6\text{H}</em>{12}\text{O}_6)</td>
</tr>
<tr>
<td>4)</td>
<td>(\text{Ca(OH)}_2)</td>
</tr>
</tbody>
</table>

Empirical formula
An empirical formula is the smallest whole # mole ratio. When given molecular formula, you can find the empirical formula by reducing. Ex: the empirical formula of \(\text{C}_6\text{H}_{12}\text{O}_6\) is \(\text{CH}_2\text{O}\) (each subscript in the molecular formula can be divided by 6). The data used to find a chemical formula is % composition or actual masses of elements. When given percent composition, assume you have 100 grams to make calculations easier. Then divide each mass by the molar mass to find the number of moles of each element in the compound. If the moles are not whole numbers, divide by the lower number of moles to reduce and express the mole ratio in the chemical formula.

Ex: Find the empirical formula composed of 40% sulfur and 60% oxygen by mass.
Convert percent to grams, assuming there are 100 grams of substance (rather than 100%).
- 40\% \text{ sulfur} = 40 \text{ g sulfur} \div 32 \text{ g/mol} = 1.25 \text{ mol S}
- 60\% \text{ oxygen} = 60 \text{ g oxygen} \div 16 \text{ g/mol} = 3.75 \text{ mol O}
Mole ratio is 1.25:3.75 \(\rightarrow\) not whole numbers so divide by the smaller value (here it is 1.25)
- \(1.25 \div 1.25 = 1 \text{ mol S}\)
- \(3.75 \div 1.25 = 3 \text{ mol O}\)
so the ratio is 1:3 of S:O
The chemical formula is \(\text{SO}_3\).
The empirical formula of a compound is the simplest whole-number ratio of atoms of the elements in a compound. It may be different from the molecular formula, which is the actual ratio of atoms in a molecule of that compound.

Molecular formula

A molecular formula specifies the actual # of atoms of each element in 1 molecule or formula unit of the substance because 2 or more substances can exist with the same % composition and empirical formula.

- Molecular formula = (empirical formula)n

Steps in determining empirical and molecular formulas:
1. Express percent by mass in grams using % composition and mass of elements.
2. Find the # of moles of each element by dividing: mass of each element ÷ molar mass. Then get the ratio of moles of elements.
3. Examine mole ratio. If there are all whole numbers you have the empirical formula. If they are not all whole numbers, multiply by the smallest factor to get whole numbers and you have the empirical formula.
4. Write the empirical formula.
5. Find the integer relating empirical and molecular formulas using: (experimental molar mass ÷ mass of empirical formula) = n.
6. Multiply subscripts by n using: (empirical formula)n.
7. Write the molecular formula.

Ex: Find the molecular formula for acetylene which has a molar mass of 26 and an empirical formula of CH. First find the mass of the empirical formula, CH is 12 + 1 = 13. Divide molecular mass by empirical mass, 26 ÷ 13 = 2, which gives the n value. Now distribute the n value to the empirical formula, 2(CH) = C₂H₂.

Practice – find the molecular formula

1) Find the molecular formula for benzene which has a molecular mass of 78 g/mol and an empirical formula of CH.

2) Find the molecular formula for the compound with a mass of 90 g/mol and empirical formula CH₂O.
The Formula for a Hydrate

Naming hydrates

A hydrate is a compound that has a specific number of water molecules bound to its atoms. In the formula, the number of water molecules associated with each formula unit is written following a dot. Ex: \( \text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} \).

Common hydrate formulas

<table>
<thead>
<tr>
<th>Prefix</th>
<th># of ( \text{H}_2\text{O} )</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mono−</td>
<td>1</td>
<td>((\text{NH}_4)_2\text{C}_2\text{O}_4 \cdot \text{H}_2\text{O})</td>
<td>Ammonium oxalate monohydrate</td>
</tr>
<tr>
<td>D1−</td>
<td>2</td>
<td>(\text{CaCl}_2 \cdot 2\text{H}_2\text{O})</td>
<td>Calcium chloride dehydrate</td>
</tr>
<tr>
<td>Tri−</td>
<td>3</td>
<td>(\text{Na}_2\text{C}_2\text{H}_3\text{O}_2 \cdot 3\text{H}_2\text{O})</td>
<td>Sodium acetate trihydrate</td>
</tr>
<tr>
<td>Tetra−</td>
<td>4</td>
<td>(\text{FePO}_4 \cdot 4\text{H}_2\text{O})</td>
<td>Iron (III) phosphate tetrahydrate</td>
</tr>
<tr>
<td>Penta−</td>
<td>5</td>
<td>(\text{CuSO}_4 \cdot 5\text{H}_2\text{O})</td>
<td>Copper (II) sulfate pentahydrate</td>
</tr>
<tr>
<td>Hexa−</td>
<td>6</td>
<td>(\text{CoCl}_2 \cdot 6\text{H}_2\text{O})</td>
<td>Cobalt (II) chloride hexahydrate</td>
</tr>
<tr>
<td>Hepta−</td>
<td>7</td>
<td>(\text{MgSO}_4 \cdot 7\text{H}_2\text{O})</td>
<td>Magnesium sulfate heptahydrate</td>
</tr>
<tr>
<td>Octa−</td>
<td>8</td>
<td>(\text{Ba(OH)}_2 \cdot 8\text{H}_2\text{O})</td>
<td>Barium hydroxide octahydrate</td>
</tr>
<tr>
<td>Nona−</td>
<td>9</td>
<td>(\text{Na}_2\text{S} \cdot 9\text{H}_2\text{O})</td>
<td>Sodium sulfide nonahydrate</td>
</tr>
<tr>
<td>Deca−</td>
<td>10</td>
<td>(\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O})</td>
<td>Sodium carbonate decahydrate</td>
</tr>
</tbody>
</table>

Analyzing a hydrate

Find the number of moles of water associated with 1 mole of the hydrate by heating the sample to drive off water of hydration. After heating, the anhydrous substance (dried/dehydrated) will have less mass. The mass of water of hydration is the difference between the mass of the hydrate and the mass of the anhydrous compound.

Ex: A 5.00 g sample of barium chloride has the formula \( \text{BaCl}_2 \cdot x \text{H}_2\text{O} \). After the sample is heated, water evaporates and the mass of the anhydrous compound is 4.26 grams. First, subtract the masses to find the mass of water; 5.00 \(-\) 4.26 = 0.74 g \( \text{H}_2\text{O} \). Then, calculate the molar mass of the compound and water; \( \text{BaCl}_2 = 208.23 \text{ g/mol} \) and \( \text{water} = 18.02 \text{ g/mol} \). Next, convert the masses to moles; \( 4.26 \text{ g} ÷ 208.23 \text{ g/mol} = 0.0205 \text{ mol BaCl}_2; 0.74 \text{ g} ÷ 18.02 \text{ g/mol} = 0.0411 \text{ mol H}_2\text{O} \). Last, calculate the mole ratio of \( \text{H}_2\text{O} \) to \( \text{BaCl}_2 \) by dividing the moles of each; 0.0411 \(÷\) 0.0205 = 2/1 which is a 2:1 ratio. Therefore, the formula is 2 mol \( \text{H}_2\text{O} \) for every 1 mol \( \text{BaCl}_2 \) making the formula \( \text{BaCl}_2 \cdot 2\text{H}_2\text{O} \).
Practice – determine the formula for the hydrate based on the experimental data

1) After a 10.00 g sample of ZnCl₂ · x H₂O is heated, 6.52 g anhydrous ZnCl₂ remains.

2) After a 5.00 g sample of NiCl₂ · x H₂O is heated, 2.72 g anhydrous NiCl₂ remains.
Chapter 11 – Stoichiometry

What is Stoichiometry?

Mole-mass relationships in chemical reactions

Stoichiometry is the study of quantitative relationships between amounts of reactants used and products formed by a reaction; it is based on the law of conservation of mass. Ex: $4\text{Fe(s)} + 3\text{O}_2 (g) \rightarrow 2\text{Fe}_2\text{O}_3 (s)$.

Mass of reactants: $4 \text{ mol Fe} \times 55.85 \text{ g} = 223.4 \text{ g Fe} \quad \text{and} \quad 3 \text{ mol O}_2 \times 32.00 \text{ g} = 96.0 \text{ g O}_2$

Total = 319.4 g

Mass of products: $2 \text{ mol Fe}_2\text{O}_3 \times 158.7 \text{ g} = 319.4 \text{ g}$

Mole ratio is the ratio between the # of moles of any 2 substances in a balanced chemical equation. You use the coefficients for each substance.

Ex: With the equation $2\text{Al (s)} + 3\text{Br}_2 \rightarrow 2\text{AlBr}_3$, there are 6 mole ratios:

- $2 \text{ mol Al} / 3 \text{ mol Br}_2$
- $3 \text{ mol Br}_2 / 2 \text{ mol Al}$
- $2 \text{ mol AlBr}_3 / 3 \text{ mol Br}_2$
- $2 \text{ mol Al} / 2 \text{ mol AlBr}_3$
- $3 \text{ mol Br}_2 / 2 \text{ mol AlBr}_3$
- $2 \text{ mol Al} / 3 \text{ mol Br}_2$

<table>
<thead>
<tr>
<th>Practice – find all mole ratios in the following chemical equations (there are 6 in each)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) $2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2$</td>
</tr>
</tbody>
</table>

Stoichiometric calculations

Using Stoichiometry

When asked to do mole-to-mole conversions first find the mole ratio. Then use the following formula:

$\text{Moles of known} \times (\text{moles of unknown in ratio} \div \text{moles of known in ration}) = \text{moles of unknown substance}$

Ex: Given the equation $2\text{K} + 2\text{H}_2\text{O} \rightarrow 2\text{KOH} + \text{H}_2$, how many moles of $\text{H}_2$ are produced when 0.04 mol of $\text{K}$ are used? The ratio of $\text{H}_2$ to $\text{K}$ is $1:2 = x:0.04$. Solve by cross multiplication to get $x = 0.02$ mol.

<table>
<thead>
<tr>
<th>Practice – answer the following questions using the equation $2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) How many moles of $\text{O}_2$ are produced when 4 mol of $\text{KClO}_3$ is used?</td>
</tr>
</tbody>
</table>
There are also mole-mass and mass-mole ratios and conversions. Steps in calculations:

1. Write the balanced equation (interpret in terms of moles).
2. Determine moles of given substances using mass-mole conversion (use inverse of molar mass).
3. Determine moles of unknown from moles of given (use mole ratio from balanced equation).
4. From moles of unknown, determine the mass of unknown using mole-mass conversion (use molar mass).

### Practice – answer the following questions using the equation \(2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2\)

<table>
<thead>
<tr>
<th>Question</th>
<th>Answer</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) If 4 mol of (\text{KClO}_3) are used, how many grams of (\text{O}_2) would be made?</td>
<td>240 g</td>
</tr>
<tr>
<td>2) If 366 g of (\text{KClO}_3) are used, how many grams of (\text{O}_2) are made?</td>
<td>90 g</td>
</tr>
<tr>
<td>3) If 74 g of (\text{KCl}) are obtained, how many moles of (\text{KClO}_3) were used?</td>
<td>1 mol</td>
</tr>
<tr>
<td>4) If 366 g of (\text{KClO}_3) are used, how many grams of (\text{KCl}) are made?</td>
<td>234 g</td>
</tr>
</tbody>
</table>

### Percent yield

How much product?

Theoretical yield is the maximum amount of product that can be produced from a given amount of reactant. Actual yield is the amount of product actually produced when the reaction is carried out in an experiment. Percent yield is the ratio of actual to theoretical yield. Formula:

\[
\text{Percent yield} = \left(\frac{\text{actual yield from experiment}}{\text{theoretical yield from calculations}}\right) \times 100.
\]

### Practice – calculate percent yield

A student performed a single replacement reaction according to the following equation:

\[2\text{Al} + 3\text{CuCl}_2 \rightarrow 2\text{AlCl}_3 + 3\text{Cu}\]

<table>
<thead>
<tr>
<th>Question</th>
<th>Answer</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) 180.0 g of (\text{Cu}) were obtained</td>
<td>86.6%</td>
</tr>
<tr>
<td>2) 237.6 g of (\text{AlCl}_3) were obtained</td>
<td>83.6%</td>
</tr>
</tbody>
</table>
Chapter 12 & 13 – States of Matter

Gases

Kinetic molecular theory

Jan Baptista Van Helmont (1577 – 1644) used the Greek word chaos (meaning without order) to describe products of reactions that had no shape or volume. Chaos became the word gas. The kinetic molecular theory (KMT) describes the behavior of gases in terms of particles of motion.

Particle size: gases consist of small particles separated by empty space. There are no significant attractive or repulsive forces. Particle motion: particles are in constant, random motion; no kinetic energy is lost during an elastic collision. Particle energy: mass and velocity determine kinetic energy; KE = ½ mv²; temperature is a measure of average kinetic energy of particles in a sample of matter.

Explaining the behavior of gases

KMT can help explain gas behavior. Low density: gases have a lot of space between particles. Compression and expansion: gas particles can be compressed together (like squeezing a pillow) and expand further apart (letting go of the squeezed pillow). Diffusion: movement of one material through another. Effusion: when gas escapes through a tiny opening. Graham’s law of effusion: the rate of effusion is inversely proportional to the square root of molar mass.

- Kinetic molecular theory (KMT) for an ideal gas states that all gas particles:
  1. are in random, constant, straight-line motion.
  2. are separated by great distances relative to their size; the volume of the gas particles is considered negligible.
  3. have no attractive forces between them.
  4. have collisions that may result in the transfer of energy between gas particles, but the total energy of the system remains constant.

- Kinetic molecular theory describes the relationships of pressure, volume, temperature, velocity, and frequency and force of collisions among gas molecules.

Gas pressure

Pressure is force per unit area. Gas particles exert pressure when they collide with walls of the container they are in. We measure air pressure with a barometer. We measure air pressure in a closed container with a manometer.

Comparison of pressure units

<table>
<thead>
<tr>
<th>Unit</th>
<th>Compared to 1 atm</th>
<th>Compared to 1 kPa</th>
</tr>
</thead>
<tbody>
<tr>
<td>Kilopascal (kPa)</td>
<td>1 atm = 101.3 kPa</td>
<td></td>
</tr>
<tr>
<td>Mm of Hg</td>
<td>1 atm = 760 mmHg</td>
<td>1 kPa = 7.501 mmHg</td>
</tr>
<tr>
<td>Torr</td>
<td>1 atm = 760 torr</td>
<td>1 kPa = 7.501 torr</td>
</tr>
<tr>
<td>Psi (or lb/in²)</td>
<td>1 atm = 14.7 psi</td>
<td>1 kPa = 0.145 psi</td>
</tr>
<tr>
<td>Atmosphere (atm)</td>
<td></td>
<td>1 kPa = 0.009869 atm</td>
</tr>
</tbody>
</table>
Dalton’s law of partial pressure states that the total pressure of a mixture of gases is equal to the sum of pressures of all the gases in the mixture. It can be used to determine the amount of gas produced by a reaction. The equation is $P_{\text{total}} = P_1 + P_2 + P_3 + \ldots$

**Forces of Attraction**

**Intermolecular forces**

Intermolecular forces are between or among particles; they are weaker than intramolecular (bonding) forces.

- Dispersion forces are weak forces resulting from temporary shifts in density of electrons in electron clouds. Dispersion forces are also called London forces and they are the weakest intermolecular force. They are the dominant force between identical nonpolar molecules.
- Dipole-dipole forces are attractions between oppositely charged regions of polar molecules.
- Hydrogen bonds are not bonds but are dipole-dipole attractions between molecules containing a hydrogen atom bonded to a small highly EN atom with at least one lone electron pair. For a hydrogen bond to form, hydrogen must bond to fluorine, oxygen or nitrogen.

- Intermolecular forces created by the unequal distribution of charge result in varying degrees of attraction between molecules. Hydrogen bonding is an example of a strong intermolecular force.
- Physical properties of substances can be explained in terms of chemical bonds and intermolecular forces. These properties include conductivity, malleability, solubility, hardness, melting point, and boiling point.

**Liquids and Solids**

**Liquids**

Density and compression: at 25°C and 1 atm liquids are much denser than gases because intermolecular forces holding particles together. Liquids can be compressed but the change in volume is smaller than in gases. Fluidity: the ability to flow; gases and liquids are both fluids but liquids are less fluid because intermolecular attractions interfere with flow.

- Viscosity: measure of resistance of a liquid to flow; stronger forces = higher viscosity.
- Viscosity and temperature: viscosity decreases with temperature increase because of the increase in average kinetic energy. Surface tension: energy required to increase surface area of a liquid by a given amount. In general, stronger attractions between particles = greater surface tension; surfactants are compounds that lower surface tension of water. Capillary action: movement of liquid up a narrow cylinder; cohesion is the attraction between identical molecules and adhesion is between different molecules.

**Solids**

Density: particles are more closely packed than liquid; most solids are more dense than most liquids. Crystalline solids: atoms, ions, or molecules are arranged in an orderly, geometric, 3-D structure; individual pieces are called crystals; a unit cell is the smallest
arrangement of connected points that can be repeated in 3 directions to form a lattice; 5 categories are atomic, molecular, covalent network, ionic and metallic.

Molecular solids – molecules are held together by dispersion forces, dipole-dipole forces or hydrogen bonds; most are not solid at room temperature, poor conductors of heat and electricity because there are no ions.

Covalent network solids – atoms like carbon and silicon that can form multiple covalent bonds are able to form covalent network solids.

Ionic solids – each ion is surrounded by ions of opposite charge. The type and ratio of ions determines the structure of lattice and shape of the crystal. They are strong but brittle. Repulsions between ions of like charges cause crystals to shatter.

Metallic solids – positive metal ions surrounded by a sea of mobile electrons. Strength of bonds between cations and electrons vary among metals which makes them have a wide range of properties. Mobile electrons make metals malleable and ductile, good conductors of heat and electricity.

Amorphous solids – particles are not arranged in a regular, repeating pattern. They form when molten material cools too quickly to allow time for crystals to form. Ex: glass, rubber, and many plastics.

- The three phases of matter (solids, liquids, and gases) have different properties

Phase Changes

Six possible transitions between phases

![Phase Change Diagram](image-url)
Practice – name the following phase changes

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1) S → L</td>
<td>3) S → G</td>
<td>5) G → L</td>
</tr>
<tr>
<td>2) G → S</td>
<td>4) L → G</td>
<td>6) L → S</td>
</tr>
</tbody>
</table>

Phase changes that require energy

Melting: solid changes to liquid; heat is the transfer of energy from an object of high temperature to an object of low temperature. The transfer is always hot to cold. The amount of energy required to melt one mole of solid depends on the strength of forces holding particles together. Melting point is the temperature at which forces holding its crystal lattice together are broken and it becomes liquid.

Vaporization: liquid changes to gas/vapor. Evaporation is not the same thing; evaporation is vaporization that occurs only at the surface of the liquid. Vapor pressure is pressure exerted by vapor over a liquid. Boiling point is the temperature at which vapor pressure of a liquid equals atmospheric pressure.

Sublimation: solid changes directly to gas without becoming a liquid first. Solid iodine and solid CO$_2$ both sublime at room temperature.

Phase changes that release energy

Freezing: reverse of melting. Freezing point is the temperature at which a liquid is converted into a crystalline solid.

Condensation: gas or vapor becomes liquid. It is the reverse of vaporization.

Deposition: gas or vapor changes directly to solid without becoming liquid first. Ex: frost on a window or snowflakes. It is the reverse of sublimation.

Practice – label the following phase changes as endothermic or exothermic

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1) Vaporization</td>
<td>4) Melting</td>
</tr>
<tr>
<td>2) Freezing</td>
<td>5) Deposition</td>
</tr>
<tr>
<td>3) Sublimation</td>
<td>6) Condensation</td>
</tr>
</tbody>
</table>

Phase diagrams

Temperature and pressure control the phase of a substance; they have opposite effects. A phase diagram is a graph of pressure versus temperature that shows which phase of a substance exists under different conditions of pressure and temperature.
Typical heating curve

- AB = solid only
- BC = melting (solid to liquid)
- CD = liquid only
- DE = vaporization (liquid to gas)
- EF = gas only

Increases in kinetic energy:
- AB, CD, and EF

No change in potential energy:
- AB, CD, and EF

No change in kinetic energy:
- BC and DE

Increases in potential energy:
- BC and DE
Typical cooling curve

The concepts of kinetic and potential energy can be used to explain physical processes that include: fusion (melting), solidification (freezing), vaporization (boiling, evaporation), condensation, sublimation, and deposition.
Chapter 12 – Gas Laws

The Gas Laws

Kinetic molecular theory
- Gas particles do not attract or repel each other.
- Gas particles are much smaller than distances between them.
- Gas particles are in constant, random motion.
- No kinetic energy is lost when gas particles collide with each other or the walls of their container.
- All gases have the same kinetic energy at a given temperature.

Nature of gases
Actual gases do not obey all the assumptions of kinetic theory. All assumptions are based on four variables: # of gas particles present, temperature, pressure and volume of gas.

Boyle’s law
Robert Boyle (1627 – 1691) was an Irish chemist. Boyle’s law states that the volume of a given amount of gas held at constant temperature varies inversely with pressure.
- Formula: \[ P_1V_1 = P_2V_2 \]

Practice – solve the following using Boyle’s law

1) A 4L balloon at 1 atm is released and the volume expands to 6L. Find the new pressure.
2) A 4L balloon at 101.3 kPa is placed under pressure of 202.6 kPa. Find the new volume.

Charles’s law
Jacques Charles (1746 – 1823) was a French physicist. Charles’s law states that the volume of a given mass of a gas is directly proportional to its Kelvin temperature at constant pressure (we do not use Celsius because you cannot divide by zero). Kelvin = 273 + Celsius.
- Formula: \( \frac{V_1}{T_1} = \frac{V_2}{T_2} \)

Practice – solve the following using Charles’s law

1) A 2L balloon at 300 K is heated to 750 K. Find the new volume.
2) A 4L balloon at 100°C deflates to 2L. Find the new temperature.
Gay-Lussac’s law
Joseph Gay-Lussac explored the relationship between temperature and pressure of a gas. Gay-Lussac’s law states that pressure of a given mass varies directly with the Kelvin temperature when volume remains constant.
- Formula: \[
\frac{P_1}{T_1} = \frac{P_2}{T_2}
\]

Practice – solve the following using Gay-Lussac’s law

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1) A gas at 202.6 kPa and 300 K is heated to 400 K. Find the new pressure.</td>
<td>2) A gas at STP is placed under a pressure of 3 atm. Find the new temperature.</td>
</tr>
</tbody>
</table>

The Combined Gas Law and Avogadro’s Principle

Combined gas law
Boyle’s law, Charles’s law and Gay-Lussac’s law can be combined into a single law known as the combined gas law. The combined gas law states the relationship among pressure, volume and temperature of a fixed amount of gas.
- Formula: \[
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}
\]

Practice – solve the following using the combined gas law

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1) A 3L canister of gas is at 1.5 atm and 200 K. It is heated to 400 K causing the volume to increase to 7L. What is the new pressure?</td>
<td>2) A 1.5L balloon at STP is placed under pressure of 148.4 kPa and temperature is lowered to 200 K. Find the new volume.</td>
</tr>
</tbody>
</table>

Avogadro’s principle
Avogadro’s principle states that equal volumes of gases at the same temperature and pressure contain equal numbers of particles. One mole of gas has \[6.02 \times 10^{23}\] particles. Molar volume for gas is the volume that one mole occupies at standard temperature and pressure (STP) which is 0°C and 1 atm (273 K and 101.3 kPa). One mole of any gas occupies 22.4 L at STP. This means that the same volume of ANY gas contains the same number of particles.
Equal volumes of different gases at the same temperature and pressure contain an equal number of particles.

The Ideal Gas Law

Ideal gas law

The # of moles is the fourth variable. For a specific amount of gas, \((PV / T) = k\) where \(k\) is a constant based on the amount of gas present \((n)\) and the ideal gas constant \((R)\) so \(k = nR\). \(R\) depends on the units used for pressure.

- Formula: \(PV = nRT\)

Numerical Values of Gas Constant \(R\)

<table>
<thead>
<tr>
<th>Units of (R)</th>
<th>(R) value</th>
<th>Units of (P)</th>
<th>(V)</th>
<th>(T)</th>
<th>(n)</th>
</tr>
</thead>
<tbody>
<tr>
<td>(L \cdot atm/mol \cdot K)</td>
<td>0.0821</td>
<td>atm</td>
<td>Liter</td>
<td>Kelvin</td>
<td>Mol</td>
</tr>
<tr>
<td>(L \cdot kPa / mol \cdot K)</td>
<td>8.314</td>
<td>kPa</td>
<td>Liter</td>
<td>Kelvin</td>
<td>Mol</td>
</tr>
<tr>
<td>(L \cdot mmHg / mol \cdot K)</td>
<td>62.4</td>
<td>mmHg</td>
<td>Liter</td>
<td>Kelvin</td>
<td>mol</td>
</tr>
</tbody>
</table>

Real versus ideal gas

An ideal gas is one whose particles take up no space and have no intermolecular forces. An ideal gas follows gas laws under all conditions of pressure and temperature. In the real world, no gas is truly ideal. Real gases deviate most from ideal gases when pressure is extremely high and temperature is very low. Real gases behave most like ideal gases under high temperature and low pressure.

- The concept of an ideal gas is a model to explain the behavior of gases. A real gas is most like an ideal gas when the real gas is at low pressure and high temperature.
Chapter 14 – Solutions

What are Solutions?

Characteristics of solutions

Solutions are homogenous mixtures with solute and solvent. Solute is the substance that dissolves and solvent is the dissolving medium. Solutions can be gas, liquid or solid, but most are liquids. Reactions can take place in aqueous solutions (reactants and products are mixed in water).

A substance that dissolves in a solvent is soluble. Ex: sugar in water. A substance that does not dissolve in a solvent is insoluble. Ex: oil and vinegar. Immiscible means substances cannot mix. Miscible means substances can mix (soluble substances are miscible).

Types and examples of solutions

<table>
<thead>
<tr>
<th>Type</th>
<th>Example</th>
<th>Solvent</th>
<th>Solute</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gas-gas</td>
<td>Air</td>
<td>Nitrogen</td>
<td>Oxygen</td>
</tr>
<tr>
<td>Gas in liquid</td>
<td>Carbonated water</td>
<td>Water</td>
<td>Carbon dioxide</td>
</tr>
<tr>
<td></td>
<td>Ocean water</td>
<td>Water</td>
<td>Oxygen</td>
</tr>
<tr>
<td>Liquid-liquid</td>
<td>Antifreeze</td>
<td>Water</td>
<td>Ethylene glycol</td>
</tr>
<tr>
<td></td>
<td>Vinegar</td>
<td>Water</td>
<td>Acetic acid</td>
</tr>
<tr>
<td>Solid in liquid</td>
<td>Ocean water</td>
<td>Water</td>
<td>Sodium chloride</td>
</tr>
<tr>
<td>Liquid in solid</td>
<td>Dental amalgam</td>
<td>Silver</td>
<td>Mercury</td>
</tr>
<tr>
<td>Solid-solid</td>
<td>Steel</td>
<td>Iron</td>
<td>Carbon</td>
</tr>
</tbody>
</table>

Solvation in aqueous solutions

Solvation is the process of surrounding solute particles with solvent particles to form a solution. Solvation in water is called hydration. Remember “like dissolves like” is the rule to see if salvation will occur. Examine bonding, polarity of particles and intermolecular forces.

Aqueous solutions of ionic compounds: attraction between dipoles and ions is greater than the attraction among ions in the crystal so ions break away from the surface. Water surrounds ions and solvated ions move into solution. Solvation continues until the entire crystal is dissolved and all ions are distributed among the solvent.

Aqueous solutions of molecular compounds: attractive forces in sugar molecules are overcome by forces between polar water molecules and polar sucrose. Oil and water do not mix because there is little attraction between nonpolar oil molecules and polar water.

Factors that affect the rate of solvation: 3 common ways to increase the number of collisions between solute and solvent are agitating the mixture (mixing, stirring, shaking), increasing surface area and increasing temperature of the solvent. Heat of solution is the overall energy change that occurs during the solution formation process.

You can determine if one substance is soluble in another by referring to Table F. There is a list of ions that will always form soluble compounds, with a few exceptions as well as a list of ions that do not form soluble compounds, with a few exceptions. Substances that do not form soluble compounds make solid precipitates.
Solubility

The maximum amount of solute that will dissolve in a given amount of solvent at a specified temperature and pressure is solubility. As long as the solvation rate is greater than the crystallization rate, solvation continues. Rate of solvation and crystallization may eventually equalize – no more solute dissolves and a state of dynamic equilibrium exists.

A saturated solution contains the maximum amount of dissolved solute for a given amount of solvent at a specific temperature and pressure. An unsaturated solution contains less dissolved solute (so more can dissolve). A supersaturated solution contains more than the maximum amount of dissolve solute for a given amount of solvent (achieved by dissolving at higher temperature and allowing solution to cool or changing pressure).

The solubility of select compounds can be found on Table G which shows solubility curves at standard pressure. When you examine the curves, locate the temperature and mass of solute dissolved in 100 grams of water for that specific compound. If the values are under the curve, the solution is unsaturated. If the values touch the curve, the solution is saturated. On certain occasions, the values will be located above the line – this means a supersaturated solution has been formed at a higher temperature and allowed to cool down.

<table>
<thead>
<tr>
<th>Practice – refer to Table G to answer the following questions</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) What is the saturation of 5 g of SO(_2) at 50°C?</td>
</tr>
<tr>
<td>2) What is the saturation of 30 g of KCl at 30°C?</td>
</tr>
<tr>
<td>3) What is the saturation of 90 g of NaNO(_3) at 20°C?</td>
</tr>
<tr>
<td>4) What mass of HCl is needed to make a saturated solution at 60°C?</td>
</tr>
<tr>
<td>5) What mass of NH(_3) is needed to make a saturated solution at 20°C?</td>
</tr>
<tr>
<td>6) At what temperature would 40 g of NaCl make a saturated solution?</td>
</tr>
<tr>
<td>7) At what temperature would 140 g of KI make a saturated solution?</td>
</tr>
<tr>
<td>8) If 120 g of KNO(_3) at 65°C is cooled to 50°C, what mass settles out if more is added?</td>
</tr>
<tr>
<td>9) If 70 g of NH(_4)Cl at 85°C is cooled to 65°C, what mass settles out if more is added?</td>
</tr>
</tbody>
</table>
Factors that affect solubility

Temperature: many substances are more soluble at higher temperatures than lower temperatures (or vice versa). As solution temperature increases, solubility of gaseous solutions decreases (think of cold soda versus warm soda – cold soda has more bubbles). The key to forming supersaturated solutions, that contain more dissolved solute than a saturated solution at the same temperature, is that some substances are more soluble at higher (or lower) temperatures.

Pressure: solubility of a gas in any solvent increases as external pressure above the solution increases (again, think of soda – a closed bottle has more bubbles and once you open it pressure is lowered so gas escapes). Henry’s law states that at a given temperature the solubility (S) of a gas in a liquid is directly proportional to the pressure (P) of the gas above the liquid. \( \frac{S_1}{P_1} = \frac{S_2}{P_2} \) or \( S_1 P_2 = S_2 P_1 \).

- A solution is a homogeneous mixture of a solute dissolved in a solvent. The solubility of a solute in a given amount of solvent is dependent on the temperature, the pressure, and the chemical natures of the solute and solvent.

Solution Concentration

- The concentration of a solution may be expressed as molarity (M), percent by volume, percent by mass, or parts per million (ppm).

Expressing concentration

Qualitative descriptions can be useful but solutions are more often described quantitatively. The description used depends on the type of solution analyzed and the reason for describing it.

Concentration ratios

<table>
<thead>
<tr>
<th>Concentration description</th>
<th>Ratio</th>
</tr>
</thead>
<tbody>
<tr>
<td>Percent by mass</td>
<td>Mass of solute ( \times 100 ) Mass of solution</td>
</tr>
<tr>
<td>Percent by volume</td>
<td>Volume of solute ( \times 100 ) Volume of solution</td>
</tr>
<tr>
<td>Molarity</td>
<td>Moles of solute Liters of solution</td>
</tr>
<tr>
<td>Parts per million</td>
<td>Grams of solute ( \times 1,000,000 ) Grams of solution</td>
</tr>
<tr>
<td>Molality</td>
<td>Moles of solute Kilogram of solution</td>
</tr>
<tr>
<td>Mole fraction</td>
<td>Moles of solute Moles of solute + moles of solvent</td>
</tr>
</tbody>
</table>
Using percent to describe concentration

Percent by mass describes a solution in which a solid is dissolved in liquid. Mass of the solution is the sum of the masses of the solute and solvent.

Percent by volume describes a solution in which both solute and solvent are liquids. Volume of the solution is the sum of the volumes of the solute and solvent.

| Practice – calculate the concentration using percent by mass or percent by volume |
|----------------------------------------|----------------------------------------|
| 1) 25 g NaCl in a 100 g solution       | 3) 76 mL O₂ in 140 mL solution          |
| 2) 25 g NaCl in 100 g water            | 4) 225 mL O₂ in 1 L solution            |

Molarity

Molarity (M) is the number of moles of solute dissolved per liter of solution. Molarity is also called molar concentration. A liter of solution with 1 mole of solute is a 1M solution, a liter of solution with 0.1 mole of solute is a 0.1M solution. You must convert to liters! 1 L = 1000 mL.

| Practice – calculate the following; you might have to convert mass to moles or vice versa |
|----------------------------------------|----------------------------------------|
| Calculate molarity                      | Given molarity, find the following     |
| 1) 2 mol NH₃ in 5 L solution            | 5) Moles in 4 L of 5 M solution        |
| 2) 3 mol KNO₃ in 8 L solution           | 6) Volume of 2 M solution with 7 mol   |
| 3) 20 g NaOH in 2 L solution            | 7) Mass of 4 L of 0.5 M KOH solution   |
| 4) 87 g NaCl in 0.5 L solution          | 8) Mass of 3 L of 0.1 M CoCl₂ solution |
Preparing molar solutions

Diluting solutions – to make solutions less concentrated. Moles of solute = molarity × liters of solution.

• Formula: $M_1V_1 = M_2V_2$ where $M_1V_1 =$ stock solution and $M_2V_2 =$ diluted solution.

To prepare a molar solution, first measure out the mass needed to get the correct molarity then add water to the amount of liters needed. Ex: To prepare 1L of a 1.5M aqueous solution of sucrose ($C_{12}H_{22}O_{11}$) first you must realize that 1.5M means you have 1.5 moles of sucrose. Molar mass of sucrose is 342 g/mol so multiply by 1.5 to get 513 g. Measure out 513 grams then add water up to 1 liter.

Parts per million

Parts per million describes very dilute concentrations of substances based on mass of the solute divided by the mass of the whole solution, multiplied by a million. It is often used when measuring levels of pollutants or contaminants.

<table>
<thead>
<tr>
<th>Practice – calculate the following</th>
<th>Find parts per million</th>
<th>Find mass of solute or solution</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) 0.025 g Pb in 125 g sample</td>
<td>2) 3.5 mg As in 100 g sample</td>
<td>5) 250 g sample contains 8 ppm O₂</td>
</tr>
<tr>
<td>3) 0.06 g Cd in 180 g sample</td>
<td>4) 0.24 g O₂ in 1680 g sample</td>
<td>6) 330 g sample contains 25 ppm Pb</td>
</tr>
<tr>
<td>7) Sample is 2.1 ppm with 0.0075 g Cd</td>
<td>8) Sample is 3.4 ppm with 0.0256 g As</td>
<td></td>
</tr>
</tbody>
</table>

Molality and mole fraction

Molality is the ratio of # of moles of solute dissolved in 1kg of solvent. The unit used is m for molal. Volume of a solution changes with temperature as it expands or contracts. Changes in volume alter molarity and since masses do not change with temperature, sometimes it is more useful to use molality as a description.
If you know the number of moles of solute and solvent you can express concentration as mole fraction. Mole fraction is the ratio of the # of moles of solute to the # of moles of solute and solvent.

Colligative Properties of Solutions

Colligative properties are physical properties of solutions that are affected by the # of particles but not the identity of the dissolved solute particles.

Electrolytes and colligative properties

Ionic compounds are called electrolytes because they dissociate in water to form a solution that conducts and electric current. Strong electrolytes produce many ions in solution and weak electrolytes produce only a few. Nonelctrolytes do not produce an electric current. Many molecular compounds dissolve in solvents but do not ionize so they do not produce current and are nonelectrolytes.

Vapor pressure lowering

The greater # of solute particles in a solvent, the lower the resulting vapor pressure. Vapor pressure lowering is due to the # of solute particles in solution.

Boiling point elevation

Boiling point elevation is the temperature difference between solution boiling point and a pure solvent’s boiling point. The greater # of solute particles in solution, the greater the boiling point elevation. For nonelectrolytes, the value of boiling point elevation is directly proportional to solution molality.

Freezing point depression

Freezing point depression is the difference in temperature between solution freezing point and freezing point of a pure solvent. For nonelectrolytes, freezing point depression is directly proportional to molality.

Osmosis and osmotic pressure

Osmosis is the diffusion of solvent particles across a semipermeable membrane from high to low solvent concentration. Osmotic pressure is the amount of additional pressure caused by water molecules that moved into the solution.

- The addition of a nonvolatile solute to a solvent causes the boiling point of the solvent to increase and the freezing point of the solvent to decrease. The greater the concentration of particles, the greater the effect.
- An electrolyte is a substance which, when dissolved in water, forms a solution capable of conducting an electric current. The ability of a solution to conduct an electric current depends on the concentration of ions.
Heterogeneous Mixtures

Suspensions
A suspension is a mixture containing particles that settle out if left undisturbed. Suspended particles are larger than solvated (dissolved) ones.

Colloids
Colloids are mixtures of intermediate sized particles (they are larger than solvated particles but smaller than suspended particles) and they do not settle out if undisturbed.

Brownian motion – jerky, random movements of colloid particles. It results from collisions of particles of the dispersion medium with dispersed particles. Collisions prevent colloid particles from settling out. Stirring electrolytes into colloid or heating it destroys the colloid and particles will settle out.

Tyndall effect – dispersed particles in a colloid are large enough to scatter light. Dilute colloids appear to be homogenous because dispersed particles are too small to be seen by the unaided eye.

Types of colloids

<table>
<thead>
<tr>
<th>Category</th>
<th>Particles</th>
<th>Medium</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solid solid</td>
<td>S</td>
<td>S</td>
<td>Colored gemstones</td>
</tr>
<tr>
<td>Solid</td>
<td>S</td>
<td>L</td>
<td>Blood</td>
</tr>
<tr>
<td>Solid emulsion</td>
<td>L</td>
<td>S</td>
<td>Butter</td>
</tr>
<tr>
<td>Emulsion</td>
<td>L</td>
<td>L</td>
<td>Milk</td>
</tr>
<tr>
<td>Solid foam</td>
<td>G</td>
<td>S</td>
<td>Marshmallow</td>
</tr>
<tr>
<td>Foam</td>
<td>G</td>
<td>L</td>
<td>Whipped cream</td>
</tr>
<tr>
<td>Aerosol</td>
<td>S</td>
<td>G</td>
<td>Smoke</td>
</tr>
<tr>
<td>Aerosol</td>
<td>L</td>
<td>G</td>
<td>Clouds</td>
</tr>
</tbody>
</table>
Chapters 2, 3, and 11 – Energy and Chemical Change

Energy

The nature of energy

Energy is the ability to do work or produce heat. Two basic forms of energy: kinetic and potential. Kinetic energy is the energy of motion. Potential energy is stored energy and is due to the composition or position of an object. Chemical systems have both forms of energy. Kinetic is directly related to constant random motion of atoms and is proportional to temperature. Potential depends on composition: type of atoms, # and type of bonds and the way atoms are rearranged.

Law of conservation of energy states that in any chemical reaction or physical process, energy can be converted from one form to another but cannot be created or destroyed.

Chemical potential energy is the energy stored in a substance because of its composition. Heat, represented by q, is energy flowing from a warmer object to a cooler object. Heat always flows from hot to cold. When a warmer object loses heat, temperature decreases. When a cooler object gains heat, temperature increases.

Measuring heat: we use the calorie (cal) or joule (J) as the unit for heat. A calorie (metric system) is the amount of heat required to raise the temperature of 1 gram of water by 1°C. The SI equivalent is joule which equals 0.2390 calories. 1 calorie = 4.184 joules.

Specific heat

Specific heat is the amount of heat required to raise the temperature of 1 gram of that substance by 1°C. Each substance has its own specific heat. Calculating heat absorbed or released: \( q = mc\Delta T \) where \( q \) = heat absorbed/released, \( m \) = mass in grams, \( c \) = specific heat and \( \Delta T \) = final temperature – initial temperature. Specific heat of water is 4.18 J/g-K. Sometimes you will have to solve for the unknown specific of a substance or a list of specific heat values will be provided – make sure to read the question before plugging in 4.18 as specific heat.

Specific heat values for selected substances

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific heat J/g-K</th>
<th>Substance</th>
<th>Specific heat J/g-K</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aluminum</td>
<td>0.900</td>
<td>Tungsten</td>
<td>0.134</td>
</tr>
<tr>
<td>Bismuth</td>
<td>0.123</td>
<td>Zinc</td>
<td>0.387</td>
</tr>
<tr>
<td>Copper</td>
<td>0.386</td>
<td>Mercury</td>
<td>0.140</td>
</tr>
<tr>
<td>Brass</td>
<td>0.380</td>
<td>Ethyl alcohol</td>
<td>2.4</td>
</tr>
<tr>
<td>Gold</td>
<td>0.126</td>
<td>Ice</td>
<td>2.05</td>
</tr>
<tr>
<td>Lead</td>
<td>0.128</td>
<td>Granite</td>
<td>0.790</td>
</tr>
<tr>
<td>Silver</td>
<td>0.233</td>
<td>Glass</td>
<td>0.84</td>
</tr>
</tbody>
</table>

- Energy can exist in different forms, such as chemical, electrical, electromagnetic, thermal, mechanical, and nuclear.
- Temperature is a measurement of the average kinetic energy of the particles in a sample of material. Temperature is not a form of energy.
### Practice – calculate the following using \( q = mc\Delta T \)

<table>
<thead>
<tr>
<th>Question</th>
<th>Answer</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) How much heat is absorbed by 10 g of water at when temperature</td>
<td>5) If 3611.52 J are used to heat 24 g of water to 78°C, find the initial temperature.</td>
</tr>
<tr>
<td>increases from 15°C to 40°C?</td>
<td></td>
</tr>
<tr>
<td>2) How much heat is released by 7 g of water when the temperature drops</td>
<td>6) What is the final temperature of 75 g of water that released 12,540 J starting at 95°C?</td>
</tr>
<tr>
<td>from 85°C to 65°C?</td>
<td></td>
</tr>
<tr>
<td>3) How much heat is released when 50 g of water is cooled from 100°C to</td>
<td>7) Find the specific heat of aluminum if it takes 450 J to heat 25 g from 70°C to 90°C.</td>
</tr>
<tr>
<td>0°C?</td>
<td></td>
</tr>
<tr>
<td>4) Find the mass of water that releases 2106.72 J when temperature changes</td>
<td>8) Find the specific heat of an unknown if 80 g loses 6512 J as temperature drops from 68°C</td>
</tr>
<tr>
<td>from 50°C to 38°C?</td>
<td>to 46°C.</td>
</tr>
</tbody>
</table>

### Heat in Chemical Reactions and Processes

#### Measuring heat

A calorimeter is an insulated device used for measuring the amount of heat absorbed or released during a chemical or physical process. Calorimeters can be used to determine specific heat of a substance.

#### Chemical energy and the universe

Virtually every chemical reaction and change in physical state either releases or absorbs heat. Thermochemistry is the study of heat changes that accompany chemical reactions and phase changes. A system is a specific part of the universe that contains the reaction or process you wish to study. Everything in the universe other than the system is considered the surroundings. The universe = the system + the surroundings.
Enthalpy (H) is the heat content of a system at constant pressure. The change in enthalpy of a reaction (heat) is called the enthalpy of a reaction ($\Delta H_{rxn}$). $\Delta H_{rxn} = H_{final} - H_{initial}$. The sign of enthalpy of a reaction indicates whether it is an exothermic or endothermic reaction. When enthalpy is negative the reaction is exothermic so $H_p < H_r$; when enthalpy is positive the reaction is endothermic so $H_p > H_r$.

- Heat is a transfer of energy (usually thermal energy) from a body of higher temperature to a body of lower temperature. Thermal energy is the energy associated with the random motion of atoms and molecules.
- Chemical and physical changes can be exothermic or endothermic

Thermochemical Equations

Writing thermochemical equations

A balanced chemical equation that includes physical states of all reactants and products as well as the energy change is called a thermochemical equation. Reference Table I shows many thermochemical equations.

Changes of state

The heat required to vaporize 1 mole of liquid is molar enthalpy (heat) of vaporization represented as $\Delta H_{vap}$. To calculate the amount of heat absorbed during vaporization or released during condensation, we use the equation $q = mH_v$ where $q =$ heat, $m =$ mass, and $H_v = 2260 \text{ J/g}$ (heat of vaporization of water). We know to use this equation when the question uses “boils”, “condenses” and/or “at 100°C”.

The heat required to melt 1 mole of solid is called molar enthalpy (heat) of fusion represented as $\Delta H_{fus}$. To calculate the amount of heat absorbed during melting or released during freezing, we use the equation $q = mH_f$ where $q =$ heat, $m =$ mass, and $H_f = 334 \text{ J/g}$ (heat of fusion of water). We know to use this equation when the question uses “melts”, “freezes” and/or “at 0°C”.

| Practice – calculate the following using equations for heat of vaporization or heat of fusion |
|---|---|
| 1) How much heat is used to melt 45 g of ice at 0°C? | 3) What mass of water releases 26,052 J of heat as it freezes at 0°C? |
| 2) How much heat is used to boil 62 g of water at 100°C? | 4) What mass of steam releases 27,120 J of heat as it condenses at 100°C? |

Calculating enthalpy change
Hess’s law

Hess’s law states that if you can add 2 or more thermochemical equations to produce a final equation for a reaction, then the sum of enthalpy changes for the individual reactions is the enthalpy change for the final reaction.

Standard enthalpy (heat) of formation

Heat of formation, \( \Delta H^\circ_f \), is the change in enthalpy that accompanies the formation of 1 mole of the compound in its standard state from its constituent elements in their standard states.

Reaction Spontaneity

Spontaneous processes

Physical or chemical change that occurs with no outside intervention is spontaneous. For many, some energy must be supplied to get the process started. Entropy (S) is a measure of the disorder or randomness of particles that make up a system. The law of disorder states that spontaneous processes always proceed in such a way that the entropy of the universe increases. Predicting changes in entropy: \( \Delta S = S_{\text{products}} - S_{\text{reactants}} \). In general, entropy increases when \( \Delta S_{\text{system}} \) is positive because \( S_{\text{products}} > S_{\text{reactants}} \) and entropy decreases when \( \Delta S_{\text{system}} \) is negative because \( S_{\text{products}} < S_{\text{reactants}} \).

Entropy changes:

1. Entropy changes associated with changes in state can be predicted. From solid to liquid or from liquid to gas, entropy increases.
2. Dissolving of gas in solvent always results in a decrease in energy.
3. Assuming no change in physical state, entropy of a system increases when the number of gaseous product particles is greater than the number of gaseous reactant particles.
4. With some exceptions, you can predict entropy change when a solid or liquid dissolves to form a solution.
5. An increase in temperature of a substance is always accompanied by an increase in random motion of its particles.

Entropy, the universe and free energy

For any spontaneous process, entropy of the universe is greater than 0. Changes in a system’s enthalpy and entropy affect \( \Delta S_{\text{universe}} \). \( \Delta S_{\text{universe}} \) tends to be positive for reactions under the following conditions:

1. A reaction or process is exothermic, \( \Delta H_{\text{system}} \) is negative.
2. Entropy of the system increases, \( \Delta S_{\text{system}} \) is positive.

Free energy is the energy available to do work. It is called Gibbs free energy represented as \( \Delta G_{\text{system}} \).

- Entropy is a measure of the randomness or disorder of a system. A system with greater disorder has greater entropy.
- Systems in nature tend to undergo changes toward lower energy and higher entropy.
Chapter 15 – Reaction Rates

A Model for Reaction Rates

Expressing reaction rates
Some chemical reactions are fast and others are slow. Average rate = Δ quantity ÷ Δ time. The equation defines average rate at which reactants produce products, which is the amount of change of reactant in a given time. Reaction rate is the change in concentration of a reactant or product per unit time, in mol / (L·s).

The collision theory
Atoms, ions, and molecules must collide to react. Collision theory explains why reactions occur and how rates can be modified. Collision theory states:
1. Reacting substances must collide.
2. Reacting substances must collide with the correct orientation.
3. Reacting substances must collide with sufficient energy to form the activated complex.

Collision theory states that a reaction is most likely to occur if reactant particles collide with the proper energy and orientation.

Orientation and the activated complex: the activated complex is a temporary and unstable arrangement of atoms that may form products or may break apart to reform reactants. It is a transition state because it is as likely to form reactants as it is to form products.

Activation energy and reaction: activation energy (E_a) is the minimum amount of energy reacting particles must have to form an activated complex and lead to reaction. Activation energy has a direct influence on reaction rate. High E_a means relatively few collisions have energy to produce activated complex and rate is low. Low E_a means more collisions have sufficient energy and rate is high.

Potential Energy Diagrams

![Exothermic Reaction](image)

![Endothermic Reaction](image)
Labeling arrows on potential energy diagrams

Potential energy of the reactants is drawn from the first straight line (on the left) to the bottom of the diagram. Potential energy of the products is drawn from the second straight line (on the right) to the bottom of the diagram. Potential energy of the activated complex is drawn from the highest part of the curve to the bottom of the diagram.

Heat of reaction is the amount of energy absorbed or released during the reaction. It can be calculated by subtracting potential energy of the products minus potential energy of the reactants. When heat of reaction is negative (higher potential energy of reactants) the reaction is exothermic. When heat of reaction is positive (higher potential energy of products) the reaction is endothermic. It is drawn between the two straight lines on the diagram.

Activation energy is the amount of energy needed to start a reaction. For exothermic reactions, activation energy is usually small because there is a lot of stored energy in the reactants. For endothermic reactions, activation energy is usually large because the reactants must absorb a lot of energy during the reaction. The arrow for activation energy is drawn from the first straight line (on the left) to the highest part of the curve.

When we read potential energy diagrams, we read them from left to right, the same way a chemical equation would proceed, such as A + B → AB (reactant side to product side). If asked about the reverse reaction, we follow the curve from right to left (in the reverse direction) so it would appear that the reaction is AB → A + B (product side to reactant side).

Practice – label the following potential energy diagrams and answer the questions

1) In terms of heat, what type of reaction is shown?
2) How much energy is required to start this reaction?
3) How much heat is released when this reaction takes place?
4) How much energy is stored in the reactants?
5) How much energy is stored in the products?
6) In terms of heat, what type of reaction is shown?
7) How much energy is required to start this reaction?
8) How much heat is absorbed when this reaction takes place?
9) How much energy is stored in the reactants?
10) How much energy is stored in the products?

- Energy released or absorbed by a chemical reaction can be represented by a potential energy diagram.
- Energy released or absorbed during a chemical reaction (heat of reaction) is equal to the difference between the potential energy of the products and the potential energy of the reactants.
- A catalyst provides an alternate reaction pathway, which has a lower activation energy than an uncatalyzed reaction.

Factors Affecting Reaction Rates

Nature of reactants
The tendency of a substance to react influences the rate of reaction. The more reactive a substance is, the faster the reaction rate.

Concentration
Reactions speed up when concentration of reacting particles is increased. The more particles there are, the more likely a collision will occur.

Surface area
Greater surface area allows molecules to collide with many more atoms per unit time.

Temperature
Generally, increasing temperature increases reaction rate. Reacting particles collide more frequently because higher temperature = more kinetic energy.
Catalysts

A catalyst is a substance that increases the rate of chemical reactions without itself being consumed in the reaction. Inhibitor is a catalyst that slows down or inhibits reaction rates. Heterogeneous catalyst exists in a physical state different than that of the reaction it catalyzes. Homogeneous catalyst exists in the same physical state. Catalysts increase the rate of reaction by lowering activation energy.

- The rate of a chemical reaction depends on several factors: temperature, concentration, nature of reactants, surface area, and the presence of a catalyst.

Chemical Equilibrium

What is equilibrium?

Reversible reactions can occur in both the forward and reverse direction. We can combine forward and reverse reactions to show balance.

Forward: \( \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \)
Equilibrium: \( \text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3 \)
Reverse: \( \text{N}_2 + 3\text{H}_2 \leftarrow 2\text{NH}_3 \)

Equilibrium means that opposing processes are in balance. Chemical equilibrium is a state in which the forward and reverse reactions balance each other because they occur at equal rates. Rate of forward = rate of reverse.

- Some chemical and physical changes can reach equilibrium.
- At equilibrium the rate of the forward reaction equals the rate of the reverse reaction.
  The measurable quantities of reactants and products remain constant at equilibrium.

Factors Affecting Chemical Equilibrium

Le Chatelier’s principle

In 1887, Le Chatelier proposed that if a stress is applied to a system at equilibrium, the system shifts in the direction that relieves the stress. A stress is any kind of change that upsets equilibrium. Le Chatelier’s principle can be used to predict how changes in concentration, volume (pressure) and temperature affect equilibrium.

Concentration

Addition or removal of a reactant or product shifts equilibrium in the direction that relieves the stress. Increased concentration of reactants shifts the reaction to the right. Decreased concentration of reactants shifts the reaction to the left. Increased concentration of products shifts the reaction to the left. Decreased concentration of products shifts the reaction to the right.

Volume

A decrease in volume increases pressure. Pressure only influences reactions that contain gases. Equilibrium will not shift if the number of moles of reactant equals the number of moles of product. If pressure on the system is increased the reaction shifts in the direction with the lower number of moles of gas. If pressure on the system is decreased the reaction shifts in the direction with the greater number of moles of gas.
Temperature
Heat can be considered as a reactant or a product. If heat is on the left side of a reaction, it is a reactant and the reaction is endothermic because heat is added. If heat is on the right side of a reaction, it is a product and the reaction is exothermic because heat is released.

Catalyst
Catalysts have no effect on equilibrium. They can make a reaction proceed faster but when a reaction is at equilibrium the rate of the forward equals the rate of the reverse so there is no shift.

- Le Chatelier’s principle can be used to predict the effect of stress (change in pressure, volume, concentration, and temperature) on a system at equilibrium.

In order to predict the direction of shift you can follow these basic steps:
1. If you are told the change in concentration or heat, use ↑ or ↓ above the substance (or heat) that is changed and put in a square.
2. Anything on the same side of the equation as the square gets an arrow opposite the square.
3. Anything on the other side of the equation gets an arrow in the same direction as what is in the square.
4. The shift is where the up arrow is (but not if it is in the square).

Practice – answer the following questions in terms of Le Chatelier’s principle

\[ \text{N}_2 (g) + 3\text{H}_2 (g) \rightleftharpoons 2\text{NH}_3 (g) + 91.8 \text{ kJ} \]

- 1) Increasing the concentration of \( \text{H}_2 \)
- 2) Increasing the temperature
- 3) Increasing pressure
- 4) Increasing the concentration of \( \text{NH}_3 \)
- 5) Decreasing the pressure
- 6) Decreasing the temperature
- 7) Decreasing concentration of \( \text{N}_2 \)
- 8) Decreasing concentration of \( \text{H}_2 \)

Predicting precipitates
You can use Reference Table F to help predict which equations will produce a precipitate. Precipitates will form if one or both products is listed as insoluble. When all substances are soluble no precipitates form.
Chapter 16 – Acids and Bases

Acids and Bases – An Introduction

Properties

Acids taste sour and bases taste bitter and feel slippery. Acids can be identified by reactions with metals. Aluminum, magnesium and zinc react to produce hydrogen gas. Metal carbonates and hydrogen carbonates react to produce carbon dioxide gas. Litmus paper can be used to distinguish between acids and bases: red litmus turns blue in base and stays red in acid; blue litmus turns red in acid and stays blue in base. Acid and base solutions conduct electricity.

Ions in solution: all water and aqueous solutions have hydrogen (H\(^+\)) and hydroxide (OH\(^-\)) ions. The amounts of these ions tell whether a solution is acidic, basic or neutral. Acidic solutions have more H\(^+\) ions, basic solutions have more OH\(^-\) ions, and neutral solutions have equal amounts of H\(^+\) and OH\(^-\) ions.

Arrhenius model: states that acids have hydrogen and ionize to make H\(^+\) ions in solution; bases have hydroxide and dissociate to make OH\(^-\) ions in solution.

- Behavior of many acids and bases can be explained by the Arrhenius theory.
- Arrhenius acids and bases are electrolytes.
- Arrhenius acids yield H\(^+\) (aq), hydrogen ion as the only positive ion in an aqueous solution. The hydrogen ion may also be written as H\(_3\)O\(^+\) (aq), hydronium ion.
- Arrhenius bases yield OH\(^-\) (aq), hydroxide ion as the only negative ion in an aqueous solution.

Bronsted-Lowry model: states that acids are H\(^+\) ion donors and bases are OH\(^-\) ion acceptors. X represents nonmetallic elements or polyatomic ions. For example, in the equation HX (aq) + H\(_2\)O (l) \(\rightleftharpoons\) H\(_3\)O\(^+\) (aq) + X\(^-\) (aq), H\(_2\)O becomes an acid by accepting H\(^+\) (acid = H\(_3\)O\(^+\)\) and acid HX becomes a base X\(^-\) by donating H\(^+\).

\[
\text{HX (aq) + H}_2\text{O (l) } \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{X}^- (aq)
\]

An acid that donates hydrogen becomes a conjugate base; a base that accepts hydrogen becomes a conjugate acid. Conjugate acid-base pair = 2 substances related to each other by the donating and accepting of a single hydrogen. Amphoteric substances can act as an acid or base (such as water). Sometimes hydronium (H\(_3\)O\(^+\)) is used instead of hydrogen (H\(^+\)) for acids.
Practice – label each part of the equations as acid, base, conjugate acid and conjugate base

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td>HF (aq) + H₂O (l) ⇌ H₃O⁺ (aq) + F⁻ (aq)</td>
</tr>
<tr>
<td>2)</td>
<td>NH₃ (aq) + H₂O (l) ⇌ NH₄⁺ (aq) + OH⁻ (aq)</td>
</tr>
</tbody>
</table>

- There are alternate acid-base theories. One theory states that an acid is an H⁺ donor and a base is an H⁺ acceptor.

Monoprotic and polyprotic acids
- Monoprotic acid can donate only 1 hydrogen ion. Ex: HCl, HF, HClO₄, HBr, CH₃COOH. Polyprotic acids can donate more than 1 hydrogen ion. Diprotic can donate 2 hydrogen ions. Ex: H₂SO₄ and H₂CO₃. Triprotic can donate 3 hydrogen ions. Ex: H₃PO₄ and H₃BO₃.
- Anhydrides – some oxides can become acids or bases by adding the elements contained in water. Oxides of nonmetallic elements (C, S, N) produce acid. Oxides of metallic elements form basic solutions.

Strength of Acids and Bases

Strength of acids
- Strong acids ionize completely, produce the maximum number of ions and are good electrical conductors. They are usually shown with → to indicate there is no reverse reaction because it ionizes completely. Ex: HCl (aq) + H₂O (l) → H₃O⁺ (aq) + Cl⁻ (aq)
- Weak acids ionize only partially in dilute aqueous solutions, produce fewer ions and cannot conduct electricity well. Weak ionization is usually indicated by ⇌ showing an equilibrium equation. Ex: HC₂H₃O₂ (aq) + H₂O (l) ⇌ H₃O⁺ (aq) + C₂H₃O₂⁻ (aq)

Strengths of bases
- Strong bases dissociate entirely into metal ions and hydroxide ions. They are usually shown with → to indicate there is no reverse reaction because it dissociates entirely. Ex: NaOH (s) → Na⁺ (aq) + OH⁻ (aq)
- Weak bases dissociate only partially in dilute aqueous solutions to form the conjugate acid of the base. Weak dissociation is usually indicated by ⇌ showing an equilibrium equation. Ex: CH₃NH₂ (aq) + H₂O (l) ⇌ CH₃NH₃⁺ (aq) + OH⁻ (aq)

Strong, weak, concentrated, dilute
- Dilute and concentrated refer to the number of acid or base molecules dissolved in a volume of solution. Weak or strong refer to the degree to which the acid or base separates into ions.
What is pH?

pH and pOH

pH is the negative logarithm of the hydrogen ion concentration.  $\text{pH} = -\log [H^+]$.

pOH is the negative logarithm of the hydroxide ion concentration.  $\text{pOH} = -\log [OH^-]$. In general chemistry, we use pH and usually do not use pOH when discussing acids and bases.

You should remember that $\text{pH} + \text{pOH} = 14.00$.

pH scale

Understanding hydrogen ion concentration

If you are asked to find pH based on hydrogen ion concentration, the exponent in the concentration tells you the pH value.  Ex: if a substance has a hydrogen ion concentration of $10^{-4}$ the pH is 4.

Many times you will be asked to compare pH in terms of hydrogen ion concentration. Since pH is a measure of the negative log of hydrogen ions there is a simple method to use when determining increase or decrease in hydrogen ion concentration.  You must remember that the more acidic a substance is, the lower the pH value and the more basic a substance is, the higher the pH value.  When pH changes occur just subtract pH values.  The number you get tells how many zeroes should be in your response.  Ex: if the pH of a substance changes from 2 to 1 it becomes more acidic.  $2 - 1 = 1$ so a pH of 1 has 10 times more hydrogen ions than a pH of 2.  If the pH changes from 2 to 3 it becomes less acidic so it will have less hydrogen ions.  $3 - 2 = 1$ but it has 10 times less (or 1/10) the amount of hydrogen ions.

Practice – determine the pH values or change in ion concentrations

| 1) 10 times more acidic than pH 5 | 7) 1/100 $[H^+]$ of pH 7 |
| 2) 10 times less $[H^+]$ than pH 5 | 8) 1/1000 $[H^+]$ of pH 4 |
| 3) 100 times more basic than pH 3 | 9) pH changes from 2 to 4 |
| 4) 100 times more acidic than pH 3 | 10) pH changes from 4 to 2 |
| 5) 1000 times more basic than pH 6 | 11) pH changes from 8 to 5 |
| 6) 1000 times more $[H^+]$ than pH 6 | 12) pH changes from 1 to 5 |

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The acidity or alkalinity of a solution can be measured by its pH value. The relative level of acidity or alkalinity of a solution can be shown by using indicators.

On the pH scale, each decrease of one unit of pH represents a tenfold increase in hydronium ion concentration.

Neutralization

The reaction between acids and bases

Neutralization reactions occur when acids and bases react together in an aqueous solution to produce salt and water. Salt is an ionic compound made of a cation from a base and an anion from an acid. Neutralization is always a double replacement reaction.

Example: HCl (aq) + NaOH (aq) → NaCl (aq) + H₂O (l).

You can examine the formula equation to get the complete ionic equation. Once you have the complete ionic equation you can cancel spectator ions to get the net ionic equation. The net ionic equation for neutralization is always H⁺ + OH⁻ → H₂O.

<table>
<thead>
<tr>
<th>Practice – complete the equations, then name and label each part (acid, base, and salt)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) Mg(OH)₂ + 2 HCl →</td>
</tr>
<tr>
<td>2) H₂SO₄ + 2KOH →</td>
</tr>
<tr>
<td>3) 2HNO₃ + Ca(OH)₂ →</td>
</tr>
<tr>
<td>4) 2NaOH + H₂CO₃ →</td>
</tr>
</tbody>
</table>

In the process of neutralization, an Arrhenius acid and an Arrhenius base react to form a salt and water.

Titration is a laboratory process in which a volume of solution of known concentration is used to determine the concentration of another solution.

Indicators

Common indicators

Indicators are chemicals that change color in the presence of acid or base to tell whether you have acid or base. Common indicators are listed in Reference Table M. To use Table M, if the substance has a pH with the first number in the pH range or lower the indicator will be the first color listed. Substances between the pH range listed will show a color change from the first color to the second color. Substances with a pH of the second number or higher will be the second color listed. Example: with methyl orange a substance with pH of 2.5 will be red, a substance with a pH of 4.0 will be a color between red and yellow (orange), and a substance with a pH of 6 will be yellow.
<table>
<thead>
<tr>
<th>Practice – use Table M to record the indicator color for each pH value</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
</tr>
<tr>
<td>Methyl orange</td>
</tr>
<tr>
<td>Bromthymol blue</td>
</tr>
<tr>
<td>Phenolphthalein</td>
</tr>
<tr>
<td>Litmus</td>
</tr>
<tr>
<td>Bromcresol green</td>
</tr>
<tr>
<td>Thymol blue</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Practice – use Table M to find the pH range based on indicator colors listed</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) bromcresol green = blue, bromthymol blue = yellow</td>
</tr>
<tr>
<td>2) bromthymol blue = blue, thymol blue = yellow</td>
</tr>
<tr>
<td>3) methyl orange = yellow, phenolphthalein = colorless</td>
</tr>
</tbody>
</table>
Chapter 17 – Redox Reactions

Oxidation and Reduction

Electron transfer and redox reactions

A reaction in which electrons are transferred from one atom to another is called an oxidation-reduction reaction (shortened to redox reaction). Oxidation numbers are found in the upper right hand corner of each element on the Periodic Table. Oxidation numbers are written with charge shown in front such as +3. Ionic charge is written after the number such as 3+.

Oxidation – loss of electrons from atoms of a substance. Electrons will be a product since they are released (lost) from the atom. Remember LEO = Lose Electrons Oxidation. Ex: Na → Na⁺ + e⁻.

Reduction – gain of electrons by atoms of a substance. Electrons will be a reactant since they are added (gained) by the atom. Remember GER = Gain Electrons Reduction. Ex: Cl₂ + 2 e⁻ → 2Cl⁻.

- An oxidation-reduction (redox) reaction involves the transfer of electrons (e⁻).
- Reduction is the gain of electrons.
- Oxidation is the loss of electrons.

Oxidizing and reducing agents

An oxidizing agent is a substance that oxidizes another substance by accepting electrons; we say the substance is reduced. A reducing agent is a substance that reduces another substance by losing electrons; we say the substance is oxidized.

Ex: The oxidized substance, K, is the reducing agent; the reduced substance, Br₂, is the oxidizing agent.

\[
\text{oxidized} \\
2K (s) + Br₂ (g) \rightarrow 2KBr (s) \\
\text{reduced}
\]

Redox and electronegativity

The more electronegative atom is treated as if it is reduced by gaining electrons and the less electronegative atom is treated as if it is oxidized by losing electrons. Refer to Table S for electronegativity values. The most electronegative atom is almost always written last, but an exception to the rule is hydrogen – it is often written last as a terminal atom.

Ex: More electronegative nitrogen is reduced and less electronegative hydrogen is oxidized.

\[
\text{oxidized} \\
N₂ + 3H₂ \rightarrow 2NH₃ \\
\text{reduced}
\]
Rules for determining oxidation numbers
1. The oxidation number for an uncombined atom (atoms not in compounds) is zero. Ex: $\text{O}_2$ and $\text{Cl}_2$ both have an oxidation number of 0.
2. The oxidation number for monatomic ions $=$ the charge on the ion. Ex: $\text{Ca}^{2+}$ has a $+2$.
3. The oxidation number of the more electronegative atom in a molecule or complex ion is the same as the charge it would have if it was an ion. Ex: In $\text{SiCl}_4$, $\text{Cl}$ is more electronegative so it has an oxidation number of $-1$ for each $\text{Cl}$ which totals to $-4$. This means that $\text{Si}$ must be $+4$.
4. The most electronegative atom fluorine ($\text{F}$) is always $-1$ when bonded to another element.
5. The oxidation number of oxygen is always $-2$ with two exceptions: if oxygen is bonded to fluorine it is $+2$ and when oxygen is in peroxide it is $-1$. Ex: In $\text{OF}_2$ oxygen is $+2$ because the 2 $\text{F}$ each has $-1$. In hydrogen peroxide $\text{H}_2\text{O}_2$, each hydrogen has $+1$ so each oxygen has $-1$.
6. The oxidation number of hydrogen in most compounds is $+1$ except when it is bonded to a less electronegative metal to form hydrides then it is $-1$. Ex: In $\text{LiH}$ the $\text{Li}$ must have $+1$ so $\text{H}$ will be $-1$.
7. Metals of groups 1, 2 and Al from group 13 form compounds in which the metal always has its oxidation number equal to the number of valence electrons.
8. The sum of oxidation numbers in neutral compounds must equal zero. You have to balance charge in the compound so it equals zero.
9. The sum of oxidation numbers in a polyatomic ion equals the charge of the ion. Common polyatomic ions are listed in Reference Table E.

<table>
<thead>
<tr>
<th>Practice – assign oxidation numbers to each of the following</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) $\text{Cu}$</td>
</tr>
<tr>
<td>2) $\text{F}_2$</td>
</tr>
<tr>
<td>3) $\text{Na}^+$</td>
</tr>
<tr>
<td>4) $\text{CO}_3^{2-}$</td>
</tr>
</tbody>
</table>

$\therefore$ Oxidation numbers (states) can be assigned to atoms and ions. Changes in oxidation numbers indicate that oxidation and reduction have occurred.
Balancing Redox Reactions

Oxidation-number method
1. Assign oxidation numbers to all atoms in the equation.
2. Identify atoms that are oxidized and reduced. When the oxidation number of an atom increases it is oxidation. When the oxidation number decreases it is reduction.
3. Determine the change in oxidation number for atoms that are oxidized and reduced.
4. Make changes in oxidation numbers equal in magnitude by adjusting coefficients in the equation.
5. If necessary, use the conventional method to balance the remainder of the equation.

Practice – assign oxidation numbers and balance the equations

1) \( \text{Fe} + \text{Cl}_2 \rightarrow \text{FeCl}_3 \)

2) \( \text{Cu} + \text{HNO}_3 \rightarrow \text{Cu(NO}_3)_2 + \text{NO}_2 + \text{H}_2\text{O} \)

Half Reactions

Identifying half reactions
One of the two parts of redox (reduction or oxidation) is a half reaction. For example, in the equation \( 2\text{Fe} + 3\text{Cl}_2 \rightarrow 3\text{FeCl}_3 \) the oxidation half reaction is: \( \text{Fe} \rightarrow \text{Fe}^{3+} + 3\text{e}^- \) because electrons are lost and the reduction half reaction is: \( \text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^- \) because electrons are gained.

- A half-reaction can be written to represent reduction.
- A half-reaction can be written to represent oxidation.

Balancing redox equations by the half reaction method
1. Write the net ionic equation and omit the spectator ions.
2. Write the oxidation and reduction half reactions.
3. Balance atoms and charges in each half reaction.
4. Adjust the coefficients so electrons lost = electrons gained.
5. Add the balanced half reaction and return spectator ions to the equation.
Practice – write half reactions for each equation; some equations must be balanced first

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1) 2Na + Cl₂ → 2NaCl</td>
<td>4) Cu + 2Ag⁺ → Cu²⁺ + 2Ag</td>
</tr>
<tr>
<td>2) Cr₂O₃ + Al → Cr + Al₂O₃</td>
<td>5) Al + O₂ → Al₂O₃</td>
</tr>
<tr>
<td>3) Cr³⁺ + Cl⁻ → Cr + Cl₂</td>
<td>6) Fe + Cd²⁺ → Fe²⁺ + Cd</td>
</tr>
</tbody>
</table>

In a redox reaction the number of electrons lost is equal to the number of electrons gained.

Electrochemistry

Redox in electrochemistry

Net ionic equations illustrate electron transfer. Two half reactions make up this process and are found in half-cells. In addition to examining the net ionic equation to determine which is oxidation and reduction you can refer to Reference Table J. First locate both substances on Table J. The one closer to the top is oxidation and the one below it is reduction. Remember when looking at Table J you see LEO then GER.

A salt bridge is the pathway constructed to allow the flow of ions from one side (half-cell) to the other. An electrochemical cell, also called a voltaic cell, converts chemical energy into electrical energy by a spontaneous redox reaction.

Practice – determine which is oxidized or reduced in each pair of elements

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1) Rb and Ba</td>
<td>3) Zn and K</td>
</tr>
<tr>
<td>2) Ca and Mn</td>
<td>4) Fe and H₂</td>
</tr>
<tr>
<td>5) Cr and Pb</td>
<td>6) Cu and Al</td>
</tr>
<tr>
<td>7) Au and Ni</td>
<td>8) Sn and Cs</td>
</tr>
</tbody>
</table>
Chemistry of voltaic cells

Electrochemical cells have 2 parts called half-cells in which separate oxidation and reduction reactions occur. The anode is the electrode where oxidation takes place and the cathode is the electrode where reduction takes place. Remember: Red Cat and An Ox. A single voltaic cell is the simplest form of a battery.
Electrolysis

Reversing redox reactions

A secondary battery can be recharged by passing a current through it in the opposite direction. Electrolysis is the use of electrical energy to force a chemical reaction to occur. An electrochemical cell in which electrolysis occurs is called an electrolytic cell.

Electrolytic cell diagram

- An electrochemical cell can be either voltaic or electrolytic. In an electrochemical cell, oxidation occurs at the anode and reduction at the cathode.
- A voltaic cell spontaneously converts chemical energy to electrical energy.
- An electrolytic cell requires electrical energy to produce chemical change. This process is known as electrolysis.
Chapter 19 – Organic Chemistry

Alkanes

Hydrocarbons

Carbon containing compounds are called organic (except carbon oxides, carbides and carbonates). The simplest organic compound is a hydrocarbon. Hydrocarbons contain only carbon and hydrogen. The simplest hydrocarbon is methane CH₄.

- Hydrocarbons are compounds that contain only carbon and hydrogen. Saturated hydrocarbons contain only single carbon-carbon bonds. Unsaturated hydrocarbons contain at least one multiple carbon-carbon bond.

Straight-chain alkanes

Alkanes are hydrocarbons that have only single bonds between atoms. Homologous series is a series of compounds that differ by a repeating unit. The alkane series can be represented by the equation CₙH₂n₊₂. To name straight-chain hydrocarbons count the number of carbons and use that prefix from Reference Table P and add –ane to the prefix. The first ten compounds of the alkane series: methane = CH₄, ethane = C₂H₆, propane = C₃H₈, butane = C₄H₁₀, pentane = C₅H₁₂, hexane = C₆H₁₄, heptanes = C₇H₁₆, octane = C₈H₁₈, nonane = C₉H₂₀, and decane = C₁₀H₂₂.

Branched-chain alkanes

Rather than the carbons forming a straight chain, in branched alkanes there are bends (branches). They have the same molecular formula but form different structural formulas (which, by definition, makes them isomers).

Straight chain versus Branched chain – both structures have the chemical formula C₄H₁₀

Butane

\[
\begin{align*}
H & \quad H \\
| & | \\
H & \quad H \\
| & | \\
H & \quad H \\
\end{align*}
\]

Isobutane

\[
\begin{align*}
H & \quad C & \quad H \\
| & | & | \\
H & \quad C & \quad C & \quad H \\
| & | & | \\
H & \quad H & \quad H & \quad H \\
\end{align*}
\]

Naming branched-chain alkanes

The longest continuous chain of carbon atoms is called the parent chain. All side branches are called substituent groups because they appear to substitute for hydrogen that would be in a straight chain. The naming process is referred to as IUPAC which stands for International Union of Pure and Applied Chemistry. Rules for naming:
1. Count the number of carbon atoms in the longest continuous chain.
2. Number each carbon in the parent chain so that the lowest number goes to the first branch on the chain.
3. Name each alkyl group substituent. To name substituents count the carbon and use the prefix for that number of carbons and add –yl instead of –ane.
4. If the same alkyl group appears more than once as a branch use another prefix to show how many there are. If there are 2 use di– if there are 3 use tri– and if there are 4 use tetra–.
5. Whenever different alkyl groups are attached to the parent chain, put the names of the groups in alphabetical order (base the alpha order on the prefix from Table P – not the prefix if there are multiple alkyl groups).
6. Write the entire name indicating the carbon number (from step 2) for each group. Use hyphens to separate numbers from words and use commas to separate numbers from other numbers.

Common alkyl groups (names of branches):
- Methyl = CH₃–
- Ethyl = CH₃CH₂–
- Propyl = CH₃CH₂CH₂–
- Butyl = CH₃CH₂CH₂CH₂–
- Isopropyl = CH₃CHCH₃

Practice – name the following branched alkanes

<table>
<thead>
<tr>
<th>1)</th>
<th>2)</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH₃</td>
<td>CH₃</td>
</tr>
<tr>
<td>CH₃–CH–CH₂–CH₃</td>
<td>CH₂–CH₃</td>
</tr>
<tr>
<td>CH₂–CH₂–CH₃</td>
<td>CH₂–CH₂–CH₂–CH₂–CH₂–CH₃</td>
</tr>
</tbody>
</table>

Practice – draw the structural formulas for the following

<table>
<thead>
<tr>
<th>1) 2-methylpropane</th>
<th>2) 4-ethyl-3,5-dimethyloctane</th>
</tr>
</thead>
</table>

Cycloalkanes

Saturated cyclic hydrocarbons form a ring. To name cycloalkanes just add cyclo– in front of the name it would have if it was a straight-chain alkane. Ex: a ring with 5 carbons would be called cyclopentane. To name substituted cycloalkanes you must first number the carbons around the ring so the lowest possible set of numbers is used for the substituents.

- Organic compounds contain carbon atoms which bond to one another in chains, rings, and networks to form a variety of structures. Organic compounds can be named using the IUPAC system.
Practice – name the following cycloalkanes

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td>CH₂–CH₂</td>
</tr>
<tr>
<td>2)</td>
<td>CH₃–CH–CH₃</td>
</tr>
</tbody>
</table>

Practice – draw the structural formulas for the following

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td>1,3-diethylcyclopentane</td>
</tr>
<tr>
<td>2)</td>
<td>1,2,4-trimethylcyclohexane</td>
</tr>
</tbody>
</table>

Alkane Properties

Properties

Physical properties: alkanes have little intermolecular attraction because they are nonpolar; boiling point and melting point are low because there are no hydrogen bonds; they are immiscible with water because water is polar and they are nonpolar.

Chemical properties: alkanes have low reactivity and there is no charge because they have nonpolar covalent bonds.

Multiple carbon-carbon bonds

Alkanes are saturated hydrocarbons and have only single bonds. They are called saturated because they have the maximum number of bonded hydrogen possible based on the number of carbon atoms. Unsaturated hydrocarbons have less than the maximum possible because they have at least 1 double or 1 triple bond. A single covalent bond has 1 pair of shared electrons. A double covalent bond has 2 pairs of shared electrons and a triple covalent bond has 3 pairs.

Alkenes and Alkynes

Alkenes

An alkene is an unsaturated hydrocarbon with 1 or more double bond. They are named by counting the number of carbon, using the prefix for that number from Table P and using the suffix –ene. The general formula for the alkene series is CₙH₂ₙ.

Alkynes

An alkyne is an unsaturated hydrocarbon with 1 or more triple bond. They are named by counting the number of carbon, using the prefix for that number from Table P and using the suffix –yne. The general formula for the alkyne series is CₙH₂ₙ₋₂.
In a multiple covalent bond, more than one pair of electrons are shared between two atoms.

Unsaturated organic compounds contain at least one double or triple bond.

Practice – name the following alkenes and alkynes

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td>( \text{CH}_3\text{CH}≡\text{CH}_2 )</td>
<td>2)</td>
</tr>
<tr>
<td>3)</td>
<td>( \text{H}≡\text{C}≡\text{H} )</td>
<td>4)</td>
</tr>
</tbody>
</table>

Practice – draw the structural formulas for the following

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td>1-butene</td>
</tr>
<tr>
<td>3)</td>
<td>2-pentyne</td>
</tr>
</tbody>
</table>

Homologous Series

Hydrocarbons can be differentiated as alkane, alkene or alkyne by examining structural formula: alkanes have single carbon-carbon bonds; alkenes have at least one double carbon-carbon bond; alkynes have at least one triple carbon-carbon bond. If structural formulas are not given, the number of carbon and hydrogen in the chemical formula can be used to determine the type of hydrocarbon by comparing it to the general formulas on Table Q. Alkanes have the general formula \( \text{C}_n\text{H}_{2n+2} \); alkenes have the general formula \( \text{C}_n\text{H}_{2n} \); alkynes have the general formula \( \text{C}_n\text{H}_{2n-2} \).
Practice – determine to which homologous series each of the following belongs

<p>| | | | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td>C₄H₈</td>
<td>4)</td>
<td>C₅H₁₂</td>
<td>7)</td>
</tr>
<tr>
<td>2)</td>
<td>C₄H₁₀</td>
<td>5)</td>
<td>C₅H₁₀</td>
<td>8)</td>
</tr>
<tr>
<td>3)</td>
<td>C₄H₆</td>
<td>6)</td>
<td>C₅H₈</td>
<td>9)</td>
</tr>
<tr>
<td>10)</td>
<td>C₆H₁₄</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>11)</td>
<td>C₆H₁₀</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>12)</td>
<td>C₆H₁₂</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Isomers

Structural isomers

Isomers are two or more compounds with the same molecular formula but different molecular structure. Structural isomers have atoms bonded in different orders. Isomers have different physical and chemical properties because structure determines properties.

Practice – draw structural isomers for each of the following

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td>[diagram]</td>
<td>2)</td>
</tr>
<tr>
<td>3)</td>
<td>3-methylpentane</td>
<td>4)</td>
</tr>
</tbody>
</table>

- Isomers of organic compounds have the same molecular formula but different structures and properties.
Substituted Hydrocarbons and Their Reactions

Functional groups

A functional group is an atom or group of atoms that always react in a certain way. Functional groups are listed with examples in Reference Table R.

<table>
<thead>
<tr>
<th>Compound</th>
<th>General Formula</th>
<th>Functional Group</th>
</tr>
</thead>
<tbody>
<tr>
<td>Halocarbon</td>
<td>( R - X )</td>
<td>Halogen</td>
</tr>
<tr>
<td>(( X ) represents any halogen)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Alcohol</td>
<td>( R - OH )</td>
<td>Hydroxyl</td>
</tr>
<tr>
<td>Ether</td>
<td>( R - O - R' )</td>
<td>Ether</td>
</tr>
<tr>
<td>Amine</td>
<td>( R' )</td>
<td>Amino</td>
</tr>
<tr>
<td></td>
<td>( R - N - R'' )</td>
<td></td>
</tr>
<tr>
<td>Aldehyde</td>
<td>( R - C = H )</td>
<td>Carbonyl</td>
</tr>
<tr>
<td>Ketone</td>
<td>( R - C = C - R' )</td>
<td>Carbonyl</td>
</tr>
<tr>
<td>Organic acid</td>
<td>( R - C = O - OH )</td>
<td>Carboxyl</td>
</tr>
<tr>
<td>Ester</td>
<td>( R - C = O - R' )</td>
<td>Ester</td>
</tr>
<tr>
<td>Amide</td>
<td>( R - C = NH )</td>
<td>Amido</td>
</tr>
</tbody>
</table>

- Organic acids, alcohols, esters, aldehydes, ketones, ethers, halides, amines, amides, and amino acids are categories of organic molecules that differ in their structures. Functional groups impart distinctive physical and chemical properties to organic compounds.

Organic compounds containing halogens

Halogens are elements in group 17 (F, Cl, Br, and I). A halocarbon is any organic compound with a halogen substituent. An alkyl halide is an organic compound containing halogen covalently bonded to an aliphatic carbon (aliphatic means carbon in an straight/open chain). Aryl halides are organic compounds containing a halogen bonded to a benzene ring or other aromatic group (aromatic means ring or cyclic).

Naming halocarbons

When naming halocarbons use the IUPAC name based on the main chain alkane with the prefix for the halogen present. We use fluoro– for fluorine, chloro– for chlorine, bromo– for bromine and iodo– for iodine.
Practice – name or draw structural formulas for the following

1) Methane

2) Chloromethane

3) Bromomethane

4) Chlorobenzene

5) 1,2-difluoropropane

6) 1-bromo-3-chloro-2-fluorobutane

7) 1,1,2-trifluoroethane

8) 1,3-dibromo-2-chlorobenzene
Alcohols, Ethers and Amines

Alcohols

When a hydroxyl group (−OH) replaces a hydrogen in a hydrocarbon it is an alcohol. Names of alcohols are based on the alkane names, drop the –e and add –ol. Ex: methane becomes methanol.

<table>
<thead>
<tr>
<th>Methane</th>
<th>Methanol</th>
</tr>
</thead>
<tbody>
<tr>
<td><img src="image" alt="Methane" /></td>
<td><img src="image" alt="Methanol" /></td>
</tr>
</tbody>
</table>

Alcohols can be categorized by the number of hydroxyl groups. When there are two hydroxyl groups the alcohol is dihydroxy and we use the name of the hydrocarbon with diol added to the end, indicating the location of hydroxyl groups by numbering carbons. When there are three hydroxyl groups the alcohol is called trihydroxy with triol added to the end, indicating the location of hydroxyl groups by numbering carbons.

<table>
<thead>
<tr>
<th>Monohydroxy</th>
<th>Dihydroxy</th>
<th>Trihydroxy</th>
</tr>
</thead>
<tbody>
<tr>
<td><img src="image" alt="2-pentanol" /></td>
<td><img src="image" alt="2,3-pentanediol" /></td>
<td><img src="image" alt="2,3,4-pentanetriol" /></td>
</tr>
</tbody>
</table>

Alcohols can also be categorized by the carbon the hydroxyl group is attached to. When the hydroxyl group is attached to a carbon that is connected to just one other carbon it is called a primary alcohol. When the hydroxyl group is attached to a carbon that is connected to two other carbon atoms it is called a secondary alcohol. When the hydroxyl group is attached to a carbon that is connected to three other carbon it is called a tertiary alcohol.

<table>
<thead>
<tr>
<th>Primary</th>
<th>Secondary</th>
<th>Tertiary</th>
</tr>
</thead>
<tbody>
<tr>
<td><img src="image" alt="Primary" /></td>
<td><img src="image" alt="Secondary" /></td>
<td><img src="image" alt="Tertiary" /></td>
</tr>
</tbody>
</table>
### Practice – name or draw structural formulas for the following

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td><img src="image1.png" alt="Image" /></td>
</tr>
<tr>
<td>2)</td>
<td><img src="image2.png" alt="Image" /></td>
</tr>
<tr>
<td>3)</td>
<td><img src="image3.png" alt="Image" /></td>
</tr>
<tr>
<td>4)</td>
<td><img src="image4.png" alt="Image" /></td>
</tr>
<tr>
<td>5)</td>
<td>3-pentanol</td>
</tr>
<tr>
<td>6)</td>
<td>1,2-butandiol</td>
</tr>
<tr>
<td>7)</td>
<td>Cyclopentanol</td>
</tr>
<tr>
<td>8)</td>
<td>4-hexanol</td>
</tr>
</tbody>
</table>

### Ethers

In ethers, oxygen is bonded to two carbon atoms. You name them by naming the two carbon groups as substituents (use the –yl ending) and then using ether at the end of the name. Ex: \( \text{CH}_3\text{OCH}_2\text{CH}_3 \) would be called methyl ethyl ether.

### Practice – name or draw structural formulas for the following

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1)</td>
<td><img src="image1.png" alt="Image" /></td>
</tr>
<tr>
<td>2)</td>
<td><img src="image2.png" alt="Image" /></td>
</tr>
<tr>
<td>3)</td>
<td>( \text{CH}_3\text{CH}_2\text{CH}_2\text{O}—\text{CH}_2\text{CH}_2\text{CH}_3 )</td>
</tr>
<tr>
<td>4)</td>
<td>( \text{CH}_3\text{CH}_2—\text{O}—\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3 )</td>
</tr>
<tr>
<td>5)</td>
<td>Cyclopentyl ether</td>
</tr>
<tr>
<td>6)</td>
<td>Methyl butyl ether</td>
</tr>
</tbody>
</table>
Amines

Amines contain nitrogen atoms bonded to carbon atoms. You name them by naming the parent chain alkane, drop the –e and add –amine to the end. Ex: CH₃CH₂NH₂ would be ethanamine.

<table>
<thead>
<tr>
<th>Practice – name or draw structural formulas for the following</th>
</tr>
</thead>
</table>
| 1) \[ \text{CH}_2\text{CH}_2\text{NH}_2 \]   | 2) \[ \text{NH}_2 \quad \text{NH}_2 \]
|                                                 | \[ \text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2 \quad \text{NH}_2 \quad \text{NH}_2 \] |
| 3) \[ \text{CH}_3\text{CH}_2\text{CH}_2\text{NH}_2 \]   | 4) \[ \text{NH}_2 \quad \text{CH}_3\text{CH}_3 \] |

| 5) Ethylamine | 6) Cyclohexylamine |

Carbonyl Compounds

Organic compounds containing the carbonyl group

A carbonyl group has an oxygen double bonded to carbon. Carbonyl groups are found in aldehydes and ketones. An aldehyde has the carbonyl group located at one end of the carbon chain. The carbonyl has carbon on one side and hydrogen on the other. Aldehydes are named by naming the parent chain alkane, drop the –e and add –al to the end. A ketone has the carbonyl located within the carbon chain so it is bonded to two carbon atoms. Ketones are named by naming the parent chain alkane, drop the –e and add –one to the end.
Organic acids

Organic acids, also called carboxylic acids, have a carboxyl group which is a carbonyl group attached to a hydroxyl group. Carboxyl groups are shown by –COOH. Organic acids are named using the name of the hydrocarbon parent chain, dropping the –e and adding –oic acid. Ex: ethane becomes ethanoic acid.

<table>
<thead>
<tr>
<th>Practice – name or draw structural formulas for the following</th>
</tr>
</thead>
</table>
| 1)  \[
\begin{array}{c}
  \text{H} \\
  \text{C} \\
  \text{H} \\
\end{array}
\] |
| 2)  \[
\begin{array}{c}
  \text{H} \\
  \text{O} \\
  \text{C} \\
\end{array}
\] |
| 3)  \[
\begin{array}{c}
  \text{H} \\
  \text{H} \\
  \text{C} \\
  \text{C} \\
  \text{H} \\
\end{array}
\] |
| 4)  \[
\begin{array}{c}
  \text{H} \\
  \text{H} \\
  \text{H} \\
  \text{H} \\
  \text{H} \\
  \text{C} \\
  \text{C} \\
  \text{C} \\
  \text{C} \\
  \text{C} \\
\end{array}
\] |
| 5) Methanal |
| 6) Methanone |

<table>
<thead>
<tr>
<th>Practice – name or draw structural formulas for the following</th>
</tr>
</thead>
</table>
| 1)  \[
\begin{array}{c}
  \text{H} \\
  \text{C} \\
  \text{C} \\
  \text{H} \\
\end{array}
\] |
| 2)  \[
\begin{array}{c}
  \text{H} \\
  \text{H} \\
  \text{C} \\
  \text{C} \\
  \text{C} \\
  \text{C} \\
  \text{C} \\
  \text{C} \\
\end{array}
\] |
| 3) Propanoic acid |
| 4) Pentanoic acid |
Organic compounds derived from organic acids

An ester has a carboxyl in which the hydrogen of the hydroxyl group has been replaced by an alkyl group. To name esters, name the alkyl group that attaches to the hydroxyl, then the –ic acid is replaced by –ate. Ex: ethanoic acid could become methyl ethanoate if a methyl group replaces the hydrogen.

\[
\text{Ethanoate group} \quad \text{Propyl group}
\]

\[
\text{CH}_3 \text{C} = \text{O} \text{CH}_2\text{CH}_2\text{CH}_3
\]

Propyl ethanoate

An amide has the hydroxyl group replaced by nitrogen. To name amides, use the name of the parent chain alkane, drop the –e and add –amide. Ex: ethane would be ethanamide.

Practice – name or draw structural formulas for the following

<table>
<thead>
<tr>
<th>1)</th>
<th>2)</th>
</tr>
</thead>
<tbody>
<tr>
<td><img src="image" alt="Methyl methanoate" /></td>
<td><img src="image" alt="Butyl ethanoate" /></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>3)</th>
<th>4)</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Methyl methanoate</strong></td>
<td><strong>Butyl ethanoate</strong></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>5)</th>
<th>6)</th>
</tr>
</thead>
<tbody>
<tr>
<td><img src="image" alt="Propanamide" /></td>
<td><img src="image" alt="Hexanamide" /></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>7)</th>
<th>8)</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Propanamide</strong></td>
<td><strong>Hexanamide</strong></td>
</tr>
</tbody>
</table>
Organic Reactions

Substitution reactions

In substitution reactions, one atom or group of atoms in a molecule is replaced by another atom or group of atoms. Halogenation occurs when hydrogen is replaced by a halogen. Once halogenated the resulting alkyl halide can undergo other substitution reactions. Reacting an alkyl halide with aqueous alkali results in halogen replacement by a hydroxyl group to form alcohol. Reacting an alkyl halide with ammonia replaces the halogen with an amino group forming an alkyl amine.

Remember: substitution, saturated, single bonds. Substitution only occurs with alkanes – saturated hydrocarbons with single bonds.

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{C} \quad \text{H} + \text{Cl}_2 \rightarrow \text{H} & \quad \text{C} \quad \text{C} \quad \text{Cl} + \text{HCl} \\
\text{H} & \quad \text{H} & \quad \text{H}
\end{align*}
\]

\[
\text{R} \quad \text{CH}_3 + \text{X}_2 \rightarrow \text{R} \quad \text{CH}_2 \text{X} + \text{HX}
\]

\[
\text{R} \quad \text{X} + \text{OH}^- \rightarrow \text{R} \quad \text{OH} + \text{X}^-
\]

Alkyl halide → Alcohol

\[
\text{R} \quad \text{X} + \text{NH}_3 \rightarrow \text{R} \quad \text{NH}_2 + \text{HX}
\]

Alkyl halide → Amine

Condensation reactions

A condensation reaction occurs when 2 smaller organic molecules combine to form a more complex molecule accompanied by a loss of a small molecule such as water. Condensation reactions are elimination reactions in which a bond is formed between 2 atoms not previously bonded.

\[
\begin{align*}
\text{H} & \quad \text{O} \\
\text{H} & \quad \text{O} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{O} & \quad \text{H} \quad \text{C} \quad \text{H}_3 & \rightarrow \text{H} & \quad \text{H} & \quad \text{O} \\
\text{O} & \quad \text{H} \quad \text{C} \quad \text{H}_3 & \quad \text{O} & \quad \text{O} & \quad \text{C} \quad \text{H}_3 & \quad \text{H}_2 \text{O}
\end{align*}
\]

Salicylic acid → Acetic acid → Acetylsalicylic acid (aspirin) → Water
Elimination reactions

Elimination reactions occur when a combination of atoms is removed from two adjacent carbon atoms forming an additional bond between carbon. Dehydrogenation eliminates 2 hydrogen atoms. Alkyl halides undergo elimination to produce an alkene and hydrogen halide. Alcohols undergo elimination by losing a hydrogen atom and hydroxyl group to form water. Dehydration is an elimination reaction in which the atoms that are removed form water.

\[ H-C-C-H \rightarrow H=C=H + H_2 \]

\[ R-CH_2-CH_2-X \rightarrow R-CH=CH_2 + HX \]

Alkyl halide          Alkene          Hydrogen halide

\[ R-CH_2-CH_2-OH \rightarrow R-CH=CH_2 + H_2O \]

Addition reactions

Addition reactions appear to be elimination reactions in reverse. Other atoms bond to each of the two atoms bonded by a double or triple bond. Common reactions with alkenes are those in which H₂O, H₂, HX or X₂ are added to an alkene. Alkynes may also undergo these addition reactions.

Addition reactions with alkenes:

Addition reactions with alkynes:

\[ R-C≡C-H + H_2 \rightarrow R-CH≡CH_2 \]  \( \text{Alkyne} \quad \text{Alkene} \)

\[ R-CH≡CH_2 + H_2 \rightarrow R-CH_2-CH_3 \]  \( \text{Alkene} \quad \text{Alkane} \)
Polymerization

Polymerization is the reaction that makes polymers. Polymers are large molecules consisting of many repeating structural units. The letter \( n \) represents the number of units in a polymer chain. A monomer is a simpler molecule from which a polymer is formed. Polymerization reactions occur when monomer units are bonded to each other in alternating sequence.

Addition polymerization occurs when all atoms present in monomers are retained in the polymer product. Condensation polymerization occurs when monomers containing at least two functional groups combine with the loss of a small by-product, usually water.

Addition polymerization:

\[
\text{H}2\text{C} = \text{C} - \text{H} + \text{H} - \text{C} - \text{C} - \text{H} \rightarrow \text{H}_2\text{C} - \text{C} - \text{C} - \text{H}_2\text{n}
\]

Ethene (ethylene)

Polyethylene

Condensation polymerization:

\[
n\text{HOOC}-(\text{CH}_2)_4-\text{COOH} + n\text{H}_2\text{N}-(\text{CH}_2)_6-\text{NH}_2 \rightarrow \left[ \text{C} - \text{C} - \text{H}_2\text{n} \right] + n\text{H}_2\text{O}
\]

Adipic acid

1,6-Diamino hexane

Nylon - 6,6

Saponification

Saponification is the process in which metallic alkali base reacts with fat or oil to produce soap.
Esterification

Esterification is a condensation reaction between alcohol and organic acid to form an ester and water.

\[
\text{RCOOH} + \text{R'OH} \rightarrow \text{RCOOR'} + \text{H}_2\text{O}
\]

- Types of organic reactions include: addition, substitution, polymerization, esterification, fermentation, saponification, and combustion.
Chapter 18 – Nuclear Chemistry

Nuclear Radiation

Nuclear chemistry versus general chemistry

Nuclear chemistry is concerned with the structure of atomic nuclei and the changes they undergo. Comparison of chemical and nuclear reactions:

<table>
<thead>
<tr>
<th>Chemical Reactions</th>
<th>Nuclear Reactions</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. They occur when bonds are formed and broken.</td>
<td>1. Occur when nuclei emit particles and/or rays.</td>
</tr>
<tr>
<td>2. Atoms are unchanged but can be rearranged.</td>
<td>2. Atoms are often converted into atoms of another element.</td>
</tr>
<tr>
<td>3. They involve only valence electrons.</td>
<td>3. They may involve protons, neutrons and electrons.</td>
</tr>
<tr>
<td>4. There are small energy changes.</td>
<td>4. There are large energy changes.</td>
</tr>
<tr>
<td>5. Reaction rate is influenced by temperature, pressure, concentration and catalysts.</td>
<td>5. Reaction rate is not normally affected by temperature, pressure, concentration or catalysts.</td>
</tr>
</tbody>
</table>

- Energy released during nuclear reactions is much greater than the energy released during chemical reactions.
- Energy released in a nuclear reaction (fission or fusion) comes from the fractional amount of mass converted into energy. Nuclear changes convert matter into energy.

Types of radiation

Radioisotopes are isotopes of atoms with unstable nuclei. They emit radiation to attain a more stable configuration in a process called radioactive decay.

Properties of Radiation

<table>
<thead>
<tr>
<th>Property</th>
<th>Alpha</th>
<th>Beta</th>
<th>Gamma</th>
</tr>
</thead>
<tbody>
<tr>
<td>Composition</td>
<td>Alpha particles</td>
<td>Beta particles</td>
<td>High energy electromagnetic radiation</td>
</tr>
<tr>
<td>Description</td>
<td>Helium</td>
<td>Electrons</td>
<td>Photons</td>
</tr>
<tr>
<td>Charge</td>
<td>2+</td>
<td>1−</td>
<td>0</td>
</tr>
<tr>
<td>Mass</td>
<td>$6.64 \times 10^{-24}$ kg</td>
<td>$9.11 \times 10^{-28}$ kg</td>
<td>0</td>
</tr>
<tr>
<td>Approximate energy</td>
<td>5 MeV</td>
<td>0.05 to 1 MeV</td>
<td>1 MeV</td>
</tr>
<tr>
<td>Relative penetrating power</td>
<td>Paper</td>
<td>Foil</td>
<td>Not fully blocked by lead or concrete</td>
</tr>
</tbody>
</table>

- Spontaneous decay can involve the release of alpha particles, beta particles, positrons and/or gamma radiation from the nucleus of an unstable isotope. These emissions differ in mass, charge, and ionizing power, and penetrating power.
Uses of common radioisotopes

Many radioisotopes are used in various industries. I-131 is used for diagnosing and treating thyroid disorders. C-14 to C-12 ratio can be used in dating once-living organisms. U-238 to Pb-206 ratio can be used in dating geological formations. Co-60 is used for treating cancer.

- There are inherent risks associated with radioactivity and the use of radioactive isotopes. Risks can include biological exposure, long-term storage and disposal, and nuclear accidents.
- Radioactive isotopes have many beneficial uses. Radioactive isotopes are used in medicine and industrial chemistry, e.g., radioactive dating, tracing chemical and biological processes, industrial measurement, nuclear power, and detection and treatment of disease.

The gold foil experiment

Ernest Rutherford (1871 – 1937) did the gold foil experiment which helped define modern atomic structure and identified alpha, beta and gamma radiation. A beam of alpha particles was aimed at a thin sheet of gold foil. The majority of the alpha particles penetrated the foil undeflected, some experienced slight deflections, a few underwent serious deflections, and others bounced back in the direction from which they came. This proved the “plum pudding” model wrong and Rutherford hypothesized that there was a tiny, dense nucleus at the center of an atom with mostly empty space surrounding the nucleus.

Radioactive Decay

Nuclear stability

Nucleons are protons and neutrons in the nucleus. Strong nuclear forces acts only on subatomic particles that are extremely close together, overcome electrostatic repulsion and hold the densely packed nucleus together.

Stability of a nucleus can be correlated to the neutron-to-proton ratio. For low atomic numbers (less than 20) atoms are most stable with a 1:1 ratio of neutrons-to-protons. As
atomic number increases, more neutrons are needed and the maximum ratio is 1.51:1. The band of stability is an area on a graph within which all stable nuclei are found. Radioactive nuclei are found outside the band of stability (above or below it). The band ends at bismuth-209. All elements with atomic number over 83 are radioactive.

Stability of isotopes is based on the ratio of neutrons and protons in its nucleus. Although most nuclei are stable, some are unstable and spontaneously decay, emitting radiation.

Types of radioactive decay
Radioactive decay is also referred to as natural transmutation because an atom of one element is converted into an atom of another element. Beta decay changes the number of protons but mass number is unchanged. Alpha decay changes both the number of protons and mass number. Positron emission and electron capture occur with nuclei that have low neutron-to-proton ratios below the band of stability. Positrons have the same mass as an electron with opposite charge. All forms of radioactive decay are represented in Reference Table O.
### Nuclear Decay Equations

<table>
<thead>
<tr>
<th>Type</th>
<th>Equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alpha decay</td>
<td>$^{226}<em>{88}\text{Ra} \rightarrow ^{222}</em>{86}\text{Rn} + ^{4}_{2}\text{He}$</td>
</tr>
<tr>
<td>Beta decay</td>
<td>$^{131}<em>{53}\text{I} \rightarrow ^{131}</em>{54}\text{Xe} + ^{0}_{-1}\beta$</td>
</tr>
<tr>
<td>Gamma decay</td>
<td>$^{238}<em>{92}\text{U} \rightarrow ^{234}</em>{90}\text{Th} + ^{4}<em>{2}\text{He} + ^{2}</em>{0}\gamma$</td>
</tr>
<tr>
<td>Positron emission</td>
<td>$^1_p \rightarrow ^1_n + ^{0}_{-1}\beta$</td>
</tr>
<tr>
<td>Electron capture</td>
<td>$^1_p + ^{-1}_0\text{e} \rightarrow ^1_0\text{n}$</td>
</tr>
</tbody>
</table>

### Practice

- Write and/or complete the following nuclear decay equations:
  
1. Beta decay of C-14
   
   $^1_p \rightarrow ^1_n + ^{0}_{-1}\beta$

2. Alpha decay of Po-210
   
   $^{141}_{61}\text{Pm} \rightarrow ^{143}_{61}\text{Nd}$

3. Positron emission of C-11
   
   $^1_p \rightarrow ^1_n$  

4. Electron capture of Rb-81
   
   $^{103}_{41}\text{Rb} + ^{-1}_0\text{e} \rightarrow ^{103}_{42}\text{Kr}$

- A change in the nucleus of an atom that converts it from one element to another is called transmutation. This can occur naturally or can be induced by the bombardment of the nucleus by high-energy particles.

### Transmutation

**Induced transmutation**

Transmutation is the conversion of an atom of one element to an atom of another element. Induced transmutation, also called artificial transmutation, involves striking nuclei with high-velocity charged particles. Induced transmutation is shown with two reactants but natural only has one because it naturally decays without human interference. Transuranium elements follow uranium in the Periodic Table (atomic number 93 and higher) and are produced in labs by induced transmutation.
Radioactive decay rates

Half-life is the time required for one half of a radioisotope’s nuclei to decay into its products. The formula for half-life is no longer found on Reference Table T. The formula is: $\text{the amount remaining} = \text{the initial amount} \times (1/2)^n$ where $n$ represents the number of half-lives that have passed. To find $n$ you can divide $t$ (elapsed time) by $T$ (duration of the half-life).

Instead of attempting the equation, treat half-life problems exactly as it sounds. Cut the initial amount in half until you get to the amount remaining. This number of “cuts” is the number of half-lives that have passed. Then multiply this number by the half-life value. Half-lives for common radioisotopes can be found in Reference Table N.

<table>
<thead>
<tr>
<th>Practice – answer the following; refer to Table N for half-life values</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) If the half-life of Fe-59 is 44.5 days, how much of a 2 mg sample remains after 133.5 days?</td>
</tr>
<tr>
<td>2) If 25 mg of strontium-90 remains after five half-lives, how much was in the original sample?</td>
</tr>
<tr>
<td>3) What is the half-life of polonium-214 if it takes 818.5 ms for 300 g to decay to 9.375 g?</td>
</tr>
<tr>
<td>4) If gallium-68 has a half-life of 68.3 minutes, how much of a 10 mg sample is left after two half-lives?</td>
</tr>
<tr>
<td>5) How long would it take for a 100 g sample of N-16 to decay to 12.5 g?</td>
</tr>
<tr>
<td>6) If a radioisotope has a half-life of 2.75 years, how much of a 10 g sample remains after four half-lives?</td>
</tr>
</tbody>
</table>

Each radioactive isotope has a specific mode and rate of decay (half-life).
Fission and Fusion of Atomic Nuclei

Nuclear reactions and energy

Einstein’s famous equation, $\Delta E = \Delta mc^2$, can show the amount of energy change in a nuclear reaction. In the equation $\Delta E$ is the energy change, $\Delta m$ is the change in mass, and $c$ is the speed of light ($3.00 \times 10^8$ m/s).

Nuclear fission

Fission is the splitting of a nucleus into many smaller fragments.

Nuclear fusion

Fusion is the combining of smaller molecules (usually hydrogen and/or helium).

- Nuclear reactions include natural and artificial transmutation, fission, and fusion.
- There are benefits and risks associated with fission and fusion reactions.
- Nuclear reactions can be represented by equations that include symbols which represent atomic nuclei (with the mass number and atomic number), subatomic particles (with mass number and charge), and/or emissions such as gamma radiation.
Questions for Regents Practice

Elements, compounds and mixtures:

1. Two substances, $A$ and $Z$, are to be identified. Substance $A$ cannot be broken down by a chemical change. Substance $Z$ can be broken down by a chemical change. What can be concluded about these substances?

   1. Both substances are elements.  
   2. Both substances are compounds.  
   3. Substance $A$ is an element and substance $Z$ is a compound.  
   4. Substance $A$ is a compound and substance $Z$ is an element.

2. Two solid samples each contain sulfur, oxygen, and sodium, only. These samples have the same color, melting point, density, and reaction with an aqueous barium chloride solution. It can be concluded that the two samples are the same

   1. compound  
   2. element  
   3. mixture  
   4. solution

3. Which of these contains only one substance?

   1. distilled water  
   2. sugar water  
   3. saltwater  
   4. rainwater

4. The accompanying table shows mass and volume data for four samples of substances at 298 K and 1 atmosphere. Which two samples could consist of the same substance?

   Masses and Volumes of Four Samples

<table>
<thead>
<tr>
<th>Sample</th>
<th>Mass (g)</th>
<th>Volume (mL)</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>30.</td>
<td>60.</td>
</tr>
<tr>
<td>B</td>
<td>40.</td>
<td>50.</td>
</tr>
<tr>
<td>C</td>
<td>45</td>
<td>90.</td>
</tr>
<tr>
<td>D</td>
<td>90.</td>
<td>120.</td>
</tr>
</tbody>
</table>

   1. $A$ and $B$  
   2. $A$ and $C$  
   3. $B$ and $C$  
   4. $C$ and $D$

5. Which sample of matter is a single substance?

   1. air  
   2. ammonia gas  
   3. hydrochloric acid  
   4. salt water

6. Tetrachloromethane, $CCl_4$, is classified as a

   1. compound because the atoms of the elements are combined in a fixed proportion  
   2. compound because the atoms of the elements are combined in a proportion that varies  
   3. mixture because the atoms of the elements are combined in a fixed proportion  
   4. mixture because the atoms of the elements are combined in a proportion that varies
7. Matter is classified as a
   1. substance, only
   2. substance or as a mixture of substances
   3. homogenous mixture, only
   4. homogenous mixture or as a heterogeneous mixture

8. A sample of unknown composition was tested in a laboratory. The sample could not be decomposed by physical or chemical means. On the basis of these results, the laboratory reported that the unknown sample was most likely
   1. a compound
   2. an element
   3. a mixture
   4. a solution

9. Base your answer on the information below.
   A student prepared two mixtures, each in a labeled beaker. Enough water at 20°C was used to make 100 milliliters of each mixture. Classify each mixture using the term "homogeneous" or the term "heterogeneous." [1]

<table>
<thead>
<tr>
<th>Information about Two Mixtures at 20°C</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Mixtue 1</strong></td>
</tr>
<tr>
<td>Composition</td>
</tr>
<tr>
<td>Student Observations</td>
</tr>
<tr>
<td></td>
</tr>
<tr>
<td>Other Data</td>
</tr>
<tr>
<td></td>
</tr>
</tbody>
</table>

**Physical and chemical properties and changes:**

1. The burning of magnesium involves a conversion of
   1. chemical energy to mechanical energy
   2. chemical energy to heat energy
   3. heat energy to chemical energy
   4. heat energy to mechanical energy

2. Which substance can not be broken down by a chemical reaction?
   1. ammonia
   2. argon
   3. methane
   4. water

3. A large sample of solid calcium sulfate is crushed into smaller pieces for testing. Which two physical properties are the same for both the large sample and one of the smaller pieces?
   1. mass and density
   2. mass and volume
   3. solubility and density
   4. solubility and volume

4. Which substance can not be broken down by a chemical change?
   1. ammonia
   2. mercury
   3. propane
   4. water
5. Which substance can *not* be decomposed by a chemical change?
   1. Ne
   2. N\textsubscript{2}O
   3. HF
   4. H\textsubscript{2}O

6. Which process represents a chemical change?
   1. melting of ice
   2. corrosion of copper
   3. evaporation of water
   4. crystallization of sugar

7. Which set of procedures and observations indicates a chemical change?
   1. Ethanol is added to an empty beaker and the ethanol eventually disappears.
   2. A solid is gently heated in a crucible and the solid slowly turns to liquid.
   3. Large crystals are crushed with a mortar and pestle and become powder.
   4. A cool, shiny metal is added to water in a beaker and rapid bubbling occurs.

8. Which two substances can *not* be broken down by chemical change?
   1. C and CuO
   2. C and Cu
   3. CO\textsubscript{2} and CuO
   4. CO\textsubscript{2} and Cu

9. Solid ZnCl\textsubscript{2} and liquid ZnCl\textsubscript{2} have different
   1. empirical formulas
   2. formula masses
   3. ion ratios
   4. physical properties

10. Which substance can *not* be decomposed by a chemical change?
    1. AlCl\textsubscript{3}
    2. H\textsubscript{2}O
    3. HI
    4. Cu

11. Which substance can *not* be decomposed by a chemical change?
    1. ammonia
    2. copper
    3. propanol
    4. water

12. Which substance can *not* be broken down by a chemical change?
    1. methane
    2. propanal
    3. tungsten
    4. water

13. Which process is a chemical change?
    1. melting of ice
    2. boiling of water
    3. subliming of ice
    4. decomposing of water

14. Two chemistry students each combine a different metal with hydrochloric acid. Student A uses zinc, and hydrogen gas is readily produced. Student B uses copper, and no hydrogen gas is produced. State one chemical reason for the different results of students A and B. [1]
15. Base your answer on the information below.

Archimedes (287-212 BC), a Greek inventor and mathematician, made several discoveries important to science today. According to a legend, Hiero, the king of Syracuse, commanded Archimedes to find out if the royal crown was made of gold, only. The king suspected that the crown consisted of a mixture of gold, tin, and copper. Archimedes measured the mass of the crown and the total amount of water displaced by the crown when it was completely submerged. He repeated the procedure using individual samples, one of gold, one of tin, and one of copper. Archimedes was able to determine that the crown was not made entirely of gold without damaging it. Identify one physical property that Archimedes used in his comparison of the metal samples.

**Density:**

1. Which element has the greatest density at STP?
   - 1. scandium
   - 2. selenium
   - 3. silicon
   - 4. sodium

2. Base your answer on the information below.
   A method used by ancient Egyptians to obtain copper metal from copper(I) sulfide ore was heating the ore in the presence of air. Later, copper was mixed with tin to produce a useful alloy called bronze. Calculate the density of a 129.5-gram sample of bronze that has a volume of 14.8 cubic centimeters. Your response must include a correct numerical setup and the calculated result. [2]

3. Base your answer on the information below.
   Archimedes (287-212 BC), a Greek inventor and mathematician, made several discoveries important to science today. According to a legend, Hiero, the king of Syracuse, commanded Archimedes to find out if the royal crown was made of gold, only. The king suspected that the crown consisted of a mixture of gold, tin, and copper. Archimedes measured the mass of the crown and the total amount of water displaced by the crown when it was completely submerged. He repeated the procedure using individual samples, one of gold, one of tin, and one of copper. Archimedes was able to determine that the crown was not made entirely of gold without damaging it. Determine the volume of a 75-gram sample of gold at STP.

4. Base your answer on the information below.
   Carbon and oxygen are examples of elements that exist in more than one form in the same phase. Graphite and diamond are two crystalline arrangements for carbon. The crystal structure of graphite is organized in layers. The bonds between carbon atoms within each layer of graphite are strong. The bonds between carbon atoms that connect different layers of graphite are weak because the shared electrons in these bonds are loosely held by the carbon atoms. The crystal structure of diamond is a strong network of atoms in which all the shared electrons are strongly held by the carbon atoms. Graphite is an electrical conductor, but diamond is not. At 25°C, graphite has a density of 2.2 g/cm³ and diamond has a density of 3.51 g/cm³. The element oxygen can exist as diatomic molecules, O₂, and as ozone, O₃. At standard pressure the boiling point of ozone is 161 K.
Calculate the volume, in cm\(^3\), of a diamond at 25°C that has a mass of 0.200 gram. Your response must include both a correct numerical setup and the calculated result. [2]

5. Base your answer on the accompanying information.
Calculate the volume of a tin block that has a mass of 95.04 grams at STP. Your response must include both a numerical setup and the calculated result. [2]

6. Base your answer on the information below.
A student prepared two mixtures, each in a labeled beaker. Enough water at 20.\(^\circ\)C was used to make 100 milliliters of each mixture.
Determine the volume of the Fe filings used to produce mixture 2. [1]

7. During a laboratory experiment, a sample of aluminum is found to have a mass of 12.50 grams and a volume of 4.6 milliliters. What is the density of this sample, expressed to the correct number of significant figures?

1. 2.717 g/mL
2. 2.72 g/mL
3. 3 g/mL
4. 2.7 g/mL

Temperature conversion:

1. The freezing point of bromine is
   1. 539°C
   2. -539°C
   3. 7°C
   4. -7°C

2. Which kelvin temperature is equal to 56°C?
   1. –329 K
   2. –217 K
   3. 217 K
   4. 329 K
3. A temperature of 37°C is equivalent to a temperature of
   1. 98.6 K
   2. 236 K
   3. 310. K
   4. 371 K

4. Base your answer on the information below.
   A weather balloon has a volume of 52.5 liters at a temperature of 295 K. The balloon is released and
   rises to an altitude where the temperature is 252 K.
   What Celsius temperature is equal to 252 K? [1]

5. Base your answer on the information below.
   A method used by ancient Egyptians to obtain copper metal from copper(I) sulfide ore was heating
   the ore in the presence of air. Later, copper was mixed with tin to produce a useful alloy called
   bronze.
   Convert the melting point of the metal obtained from copper(I) sulfide ore to degrees Celsius. [1]

6. Base your answer on the accompanying information.
   Convert the boiling point of hydrogen chloride at standard pressure to kelvins. [1]

<table>
<thead>
<tr>
<th>Compound</th>
<th>Boiling Point (°C)</th>
<th>Solubility in 100. Grams of H₂O at 20.°C (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>ammonia</td>
<td>−33.2</td>
<td>56</td>
</tr>
<tr>
<td>methane</td>
<td>−161.5</td>
<td>0.002</td>
</tr>
<tr>
<td>hydrogen chloride</td>
<td>−84.9</td>
<td>72</td>
</tr>
</tbody>
</table>

7. What kelvin temperature is equal to 15°C? [1]

8. Base your answer on the information shown, which describes the smelting of iron ore, and on your
   knowledge of chemistry.
   In the smelting of iron ore, Fe₂O₃ is reduced in a blast furnace at high temperature by a reaction with
   carbon monoxide. Crushed limestone, CaCO₃, is also added to the mixture to remove impurities in the
   ore. The carbon monoxide is formed by the oxidation of carbon (coke), as shown in the reaction
   shown:
   
   \[ 2 \text{C} + \text{O}_2 \rightarrow 2 \text{CO} + \text{energy} \]

   Liquid iron flows from the bottom of the blast furnace and is processed into different allays of iron.
   Convert the melting point of iron metal to degrees Celsius. [1]

**Unit conversion:**

1. Which quantity of heat is equal to 200. joules?
   1. 20.0 kJ
   2. 2.00 kJ
   3. 0.200 kJ
   4. 0.0200 kJ
2. Base your answer on the information below.

Arsenic is often obtained by heating the ore arsenopyrite, FeAsS. The decomposition of FeAsS is represented by the balanced equation shown.

In the solid phase, arsenic occurs in two forms. One form, yellow arsenic, has a density of 1.97 g/cm$^3$ at STP. The other form, gray arsenic, has a density of 5.78 g/cm$^3$ at STP. When arsenic is heated rapidly in air, arsenic(III) oxide is formed.

Although arsenic is toxic, it is needed by the human body in very small amounts. The body of a healthy human adult contains approximately 5 milligrams of arsenic.

Convert the mass of arsenic found in the body of a healthy human adult to grams. [1]

\[
\text{FeAsS(s)} \xrightarrow{\text{heat}} \text{FeS(s)} + \text{As(g)}
\]

**Percent error:**

1. A student measures the mass and volume of a piece of aluminum. The measurements are 25.6 grams and 9.1 cubic centimeters. The student calculates the density of the aluminum. What is the percent error of the student's calculated density of aluminum?
   1. 1%
   2. 2%
   3. 3%
   4. 4%

2. Based on data collected during a laboratory investigation, a student determined an experimental value of 322 joules per gram for the heat of fusion of H$_2$O. Calculate the student's percent error. Your response must include a correct numerical setup and the calculated result. [2]

3. Base your answer on the information below.

The accepted values for the atomic mass and percent natural abundance of each naturally occurring isotope of silicon are given in the accompanying data table.

A scientist calculated the percent natural abundance of Si-30 in a sample to be 3.29%. Determine the percent error for this value. [1]

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Atomic Mass (atomic mass units)</th>
<th>Percent Natural Abundance (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Si-28</td>
<td>27.98</td>
<td>92.22</td>
</tr>
<tr>
<td>Si-29</td>
<td>28.98</td>
<td>4.69</td>
</tr>
<tr>
<td>Si-30</td>
<td>29.97</td>
<td>3.09</td>
</tr>
</tbody>
</table>
4. A student used a balance and a graduated cylinder to collect the following data shown:
   a) If the accepted value is 6.93 grams per milliliter, calculate the percent error. [1]
   b) What error is introduced if the volume of the sample is determined first? [1]

<table>
<thead>
<tr>
<th>Sample mass</th>
<th>10.23 g</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume of water</td>
<td>20.0 mL</td>
</tr>
<tr>
<td>Volume of water and sample</td>
<td>21.5 mL</td>
</tr>
</tbody>
</table>

5. A glass tube is filled with hydrogen gas at low pressure. An electric current is passed through the gas, causing it to emit light. This light is passed through a prism to separate the light into the bright, colored lines of hydrogen's visible spectrum. Each colored line corresponds to a particular wavelength of light. One of hydrogen's spectral lines is red light with a wavelength of 656 nanometers. Tubes filled with other gases produce different bright-line spectra that are characteristic of each kind of gas. These spectra have been observed and recorded.

A student measured the wavelength of hydrogen's visible red spectral line to be 647 nanometers. *On a separate piece of paper, show a correct numerical setup for calculating the student's percent error.*

**Significant figures:**

1. Expressed to the correct number of significant figures, the sum of two masses is 445.2 grams. Which two masses produce this answer?
   1. 210.10 g + 235.100 g
   2. 210.100 g + 235.10 g
   3. 210.1 g + 235.1 g
   4. 210.10 g + 235.10 g

2. A sample of an element has a mass of 34.261 grams and a volume of 3.8 cubic centimeters. To which number of significant figures should the calculated density of the sample be expressed?
   1. 5
   2. 2
   3. 3
   4. 4

3. A student calculates the density of an unknown solid. The mass is 10.04 grams, and the volume is 8.21 cubic centimeters. How many significant figures should appear in the final answer?
   1. 1
   2. 2
   3. 3
   4. 4

4. The atomic mass of Cu-63 is expressed to what number of significant figures? [1]

   **Naturally Occurring Isotopes of Copper**

<table>
<thead>
<tr>
<th>Isotope Notation</th>
<th>Percent Natural Abundance (%)</th>
<th>Atomic Mass (atomic mass units, u)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cu-63</td>
<td>69.17</td>
<td>62.930</td>
</tr>
<tr>
<td>Cu-65</td>
<td>30.83</td>
<td>64.928</td>
</tr>
</tbody>
</table>
5. Base your answer on the information below.
Identify the total number of significant figures recorded in the calculated mass of CuSO₄•5H₂O(s). [1]

Data and calculation before heating:

\[
\begin{align*}
\text{mass of CuSO₄•5H₂O(s) and crucible} & = 21.37 \text{ g} \\
\text{mass of crucible} & = 19.24 \text{ g} \\
\text{mass of CuSO₄•5H₂O(s)} & = 2.13 \text{ g}
\end{align*}
\]

Data and calculation after heating to a constant mass:

\[
\begin{align*}
\text{mass of CuSO₄(s) and crucible} & = 20.61 \text{ g} \\
\text{mass of crucible} & = 19.24 \text{ g} \\
\text{mass of CuSO₄(s)} & = 1.37 \text{ g}
\end{align*}
\]

Calculation to determine the mass of water:

\[
\begin{align*}
\text{mass of CuSO₄•5H₂O(s)} & = 2.13 \text{ g} \\
\text{mass of CuSO₄(s)} & = 1.37 \text{ g} \\
\text{mass of H₂O(g)} & = 0.76 \text{ g}
\end{align*}
\]

6. In a titration, 3.00 M NaOH(aq) was added to an Erlenmeyer flask containing 25.00 milliliters of HCl(aq) and three drops of phenolphthalein until one drop of the NaOH(aq) turned the solution a light-pink color. The following data were collected by a student performing this titration.

Initial NaOH(aq) buret reading: 14.45 milliliters
Final NaOH(aq) buret reading: 32.66 milliliters

Based on the data given, what is the correct number of significant figures that should be shown in the molarity of the HCl(aq)?

7. In performing a titration, a student adds three drops of phenolphthalein to a flask containing 25.00 milliliters of HCl(aq). Using a buret, the student slowly adds 0.150 M NaOH(aq) to the flask until one drop causes the indicator to turn light pink. The student determines that a total volume of 20.20 milliliters of NaOH(aq) was used in this titration.

The concentration of the NaOH(aq) used in the titration is expressed to what number of significant figures? [1]

8. A student titrates 60.0 mL of HNO₃(aq) with 0.30 M NaOH(aq). Phenolphthalein is used as the indicator. After adding 42.2 mL of NaOH(aq), a color change remains for 25 seconds, and the student stops the titration.

According to the data, how many significant figures should be present in the calculated molarity of the HNO₃(aq)? [1]
Phase diagrams:

1. Which diagram in the image shown best represents a gas in a closed container?

2. Base your answer on the information shown: . . . (see image)
   a Draw a particle model that shows at least six molecules of nitrogen gas. [1]
   b Draw a particle model that shows at least six molecules of liquid nitrogen. [1]
   c Describe, in terms of particle arrangement, the difference between nitrogen gas and liquid nitrogen. [1]
   d Good models should reflect the true nature of the concept being represented. What is a limitation of two-dimensional models? [1]

   ○ represents one molecule of nitrogen

3. Heat is added to a sample of liquid water, starting at 80.°C, until the entire sample is a gas at 120.°C. This process, occurring at standard pressure, is represented by the balanced equation below.

   \[ \text{H}_2\text{O(l)} + \text{heat} \rightarrow \text{H}_2\text{O(g)} \]

   In the box provided or on a separate piece of paper, using the key, draw a particle diagram to represent at least five molecules of the product of this physical change at 120.°C. [2]

<table>
<thead>
<tr>
<th>Key</th>
</tr>
</thead>
<tbody>
<tr>
<td>● = atom of hydrogen</td>
</tr>
<tr>
<td>○ = atom of oxygen</td>
</tr>
</tbody>
</table>

4. Base your answer on the information below.

   A portable propane-fueled lantern contains a mesh silk bag coated with metal hydroxides. The primary metal hydroxide is yttrium hydroxide. When the silk bag is installed, it is ignited and burned
away, leaving the metal hydroxide coating. The coating forms metal oxides that glow brightly when heated to a high temperature. During a test, a propane lantern is operated for three hours and consumes 5.0 moles of propane from the lantern's tank. The balanced equation below represents the combustion of propane.

\[ \text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O} + \text{energy} \]

At standard pressure, the boiling point of propane is 231 K. In the box on the answer sheet or on a separate piece of paper, draw a particle diagram to represent the phase of the propane as it leaves the tank at 294 K. Your response must include at least six molecules. [1]

5. Base your answer on the information below.

The equilibrium equation shown is related to the manufacture of a bleaching solution. In this equation, \( \text{Cl}^-\) (aq) means that chloride ions are surrounded by water molecules.

In the space on the answer sheet, or, if taken online, on a separate piece of paper, use the key to draw two water molecules in the box, showing the correct orientation of each water molecule toward the chloride ion. [1]

\[ \text{Cl}_2(g) + 2\text{OH}^-(aq) \rightleftharpoons \text{OCl}^-(aq) + \text{Cl}^-(aq) + \text{H}_2\text{O}(\ell) \]

6. Base your answer on the information below.

The temperature of a sample of a substance is increased from 20.°C to 160.°C as the sample absorbs heat at a constant rate of 15 kilojoules per minute at standard pressure. The accompanying graph represents the relationship between temperature and time as the sample is heated.

In the space provided, or, if taken online, on a separate piece of paper, use the key to draw at least
nine particles in the box, showing the correct particle arrangement of this sample during the first minute of heating. [1]

7. Given the heating curve (see image) where substance $X$ starts as a solid below its melting point and is heated uniformly:
Using dots to represent particles of substance $X$, draw at least five particles as they would appear in the substance at point $F$. Use the box provided (print out this page). [1]
8. Base your answer on the information below.
Given the balanced equation for dissolving NH₄Cl(s) in water (see accompanying diagram):
Using the key shown in the accompanying diagram, draw at least two water molecules in the box (or on a separate piece of paper), showing the correct orientation of each water molecule when it is near the Cl⁻ ion in the aqueous solution.

\[ \text{NH}_4\text{Cl(s)} \rightarrow \text{H}_2\text{O} \rightarrow \text{NH}_4^+(aq) + \text{Cl}^-(aq) \]

<table>
<thead>
<tr>
<th>Key</th>
</tr>
</thead>
<tbody>
<tr>
<td>●   = Hydrogen atom</td>
</tr>
<tr>
<td>○   = Oxygen atom</td>
</tr>
<tr>
<td>○●  = Water molecule</td>
</tr>
</tbody>
</table>

Particle diagrams:

1. Using the given diagram: Which diagram represents a mixture?

   ● = particle X
   ○ = particle Y

(1) (2) (3) (4)
2. Given the simple representations for atoms of two elements in the accompanying diagram: Which particle diagram represents molecules of only one compound in the gaseous phase?

- ○ = an atom of an element
- ● = an atom of a different element

3. Which particle diagram represents a sample of one compound, only?
4. Which particle model diagram represents only one compound composed of elements X and Z?

<table>
<thead>
<tr>
<th>Key</th>
</tr>
</thead>
<tbody>
<tr>
<td>● = atom of element X</td>
</tr>
<tr>
<td>○ = atom of element Z</td>
</tr>
</tbody>
</table>

(1)  
(3)  
(2)  
(4)

5. Which two particle diagrams represent mixtures of diatomic elements?

<table>
<thead>
<tr>
<th>Key</th>
</tr>
</thead>
<tbody>
<tr>
<td>○ = atom of one element</td>
</tr>
<tr>
<td>● = atom of another element</td>
</tr>
</tbody>
</table>

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

6. Which diagram represents a physical change, only?

<table>
<thead>
<tr>
<th>Key</th>
</tr>
</thead>
<tbody>
<tr>
<td>● = an atom of an element</td>
</tr>
<tr>
<td>○ = an atom of a different element</td>
</tr>
</tbody>
</table>

(1)  
(3)  
(2)  
(4)
7. Which particle diagram shown represents a mixture of element X and element Z, only?

8. Given the balanced particle-diagram equation: Which statement describes the type of change and the chemical properties of the product and reactants?

9. Base your answer on the particle diagrams (see image). Samples A, B, and C contain molecules at STP. Explain why the average kinetic energy of sample B is equal to the average kinetic energy of sample C. [1]

10. On a field trip, Student X and Student Y collected two rock samples. Analysis revealed that both rocks contained lead and sulfur. One rock contained a certain percentage of lead and sulfur by mass, and the other rock contained a different percentage of lead and sulfur by mass. Student X stated that the rocks contained two different mixtures of lead and sulfur. Student Y stated that the rocks contained two different compounds of lead and sulfur. Their teacher stated that both students could be correct. Draw particle diagrams in each of the rock diagrams provided to show how Student X’s and Student Y’s explanations could both be correct. Use the symbols in the key provided to sketch lead and sulfur atoms. [2]
A method used by ancient Egyptians to obtain copper metal from copper(I) sulfide ore was heating the ore in the presence of air. Later, copper was mixed with tin to produce a useful alloy called bronze.

A 133.8-gram sample of bronze was 10.3% tin by mass. Determine the total mass of tin in the sample.

2. A hydrate is a compound that has water molecules within its crystal structure. The formula for the hydrate CuSO₄•5H₂O(s) shows that there are five moles of water for every one mole of CuSO₄(s). When CuSO₄•5H₂O(s) is heated, the water within the crystals is released, as represented by the balanced equation below.

\[
\text{CuSO}_4\cdot5\text{H}_2\text{O}(s) \rightarrow \text{CuSO}_4(s) + 5\text{H}_2\text{O}(g)
\]

A student first masses an empty crucible (a heat-resistant container). The student then masses the crucible containing a sample of CuSO₄•5H₂O(s). The student repeatedly heats and masses the crucible and its contents until the mass is constant. The student's recorded experimental data and calculations are shown in the accompanying diagram.

In the space on the answer sheet, use the student's data to show a correct numerical setup for calculating the percent composition by mass of water in the hydrate. [1]
Atomic structure and subatomic particles:

1. Which list of particles is in order of increasing mass?
   (1) proton → electron → alpha particle
   (2) proton → alpha particle → electron
   (3) electron → proton → alpha particle
   (4) alpha particle → electron → proton

2. Which statement is true about the charges assigned to an electron and a proton?
   1. Both an electron and a proton are positive.  3. An electron is negative and a proton is positive.
   2. An electron is positive and a proton is negative.  4. Both an electron and a proton are negative.

3. In the wave-mechanical model, an orbital is a region of space in an atom where there is
   1. a high probability of finding an electron  3. a circular path in which electrons are found
   2. a high probability of finding a neutron  4. a circular path in which neutrons are found

4. How many electrons are contained in an Au$^{3+}$ ion?
   1. 76  3. 82
   2. 79  4. 197

5. In the modern wave-mechanical model of the atom, the orbitals are regions of the most probable location of
   1. protons  3. electrons
   2. neutrons  4. positrons

6. Compared to a proton, an electron has
   1. a greater quantity of charge and the same sign  3. the same quantity of charge and the same sign
   2. a greater quantity of charge and the opposite sign  4. the same quantity of charge and the opposite sign

7. Given the accompanying table that shows students' examples of proposed models of the atom: Which model correctly describes the locations of protons and electrons in the wave-mechanical model of the atom?
8. According to the wave-mechanical model of the atom, electrons in an atom
   1. travel in defined circles
   2. are most likely found in an excited state
   3. have a positive charge
   4. are located in orbitals outside the nucleus

9. What is the total charge of the nucleus of a carbon atom?
   1. -6
   2. 0
   3. +6
   4. +12

10. Which two particles each have a mass approximately equal to one atomic mass unit?
    1. electron and neutron
    2. electron and positron
    3. proton and electron
    4. proton and neutron

11. What is the total number of electrons in an atom of potassium?
    1. 18
    2. 19
    3. 20
    4. 39

12. A proton has a charge that is opposite the charge of
    1. an alpha particle
    2. a neutron
    3. an electron
    4. a positron

13. The wave-mechanical model of the atom is required to explain the
    1. mass number and atomic number of an atom
    2. organization of atoms in a crystal
    3. radioactive nature of some atoms
    4. spectra of elements with multielectron atoms

14. The accompanying diagram represents the nucleus of an atom.
What are the atomic number and mass number of this atom?
1. The atomic number is 9 and the mass number is 19.  
2. The atomic number is 9 and the mass number is 20.  
3. The atomic number is 11 and the mass number is 19.  
4. The atomic number is 11 and the mass number is 20.

15. Which subatomic particle is negatively charged?  
   1. electron  
   2. neutron  
   3. positron  
   4. proton

16. Which subatomic particles are located in the nucleus of a carbon atom?  
   1. protons, only  
   2. neutrons, only  
   3. protons and neutrons  
   4. protons and electrons

17. Which part of a helium atom is positively charged?  
   1. electron  
   2. neutron  
   3. nucleus  
   4. orbital

18. The mass of a proton is approximately equal to the mass of  
   1. an alpha particle  
   2. an electron  
   3. a neutron  
   4. a positron

19. A neutron has a charge of  
   1. +1  
   2. +2  
   3. 0  
   4. -1

20. In the electron cloud model of the atom, an orbital is defined as the most probable  
   1. charge of an electron  
   2. conductivity of an electron  
   3. location of an electron  
   4. mass of an electron

21. Subatomic particles can usually pass undeflected through an atom because the volume of an atom is composed of  
   1. an uncharged nucleus  
   2. largely empty space  
   3. neutrons  
   4. protons

22. An electron has a charge of  
   1. -1 and the same mass as a proton  
   2. +1 and the same mass as a proton  
   3. -1 and a smaller mass than a proton  
   4. +1 and a smaller mass than a proton
23. Which statement is true about a proton and an electron?
   1. They have the same masses and the same charges.
   2. They have the same masses and different charges.
   3. They have different masses and the same charges.
   4. They have different masses and different charges.

24. In which pair do the particles have approximately the same mass?
   1. proton and electron
   2. proton and neutron
   3. neutron and electron
   4. neutron and beta particle

25. Note: This question may require the use of the Reference Tables for Physical Setting/Chemistry.
Which symbol represents a particle with a total of 10 electrons?
   1. N
   2. N
   3. Al
   4. Al

26. Which statement correctly describes the charge of the nucleus and the charge of the electron cloud of an atom?
   1. The nucleus is positive and the electron cloud is positive.
   2. The nucleus is positive and the electron cloud is negative.
   3. The nucleus is negative and the electron cloud is positive.
   4. The nucleus is negative and the electron cloud is negative.

27. An atom is electrically neutral because the
   1. number of protons equals the number of electrons
   2. number of protons equals the number of neutrons
   3. ratio of the number of neutrons to the number of electrons is 1:1
   4. ratio of the number of neutrons to the number of protons is 2:1

28. Which two particles make up most of the mass of a hydrogen-2 atom?
   1. electron and neutron
   2. electron and proton
   3. proton and neutron
   4. proton and positron

29. In the wave-mechanical model of the atom, orbitals are regions of the most probable locations of
   1. protons
   2. positrons
   3. neutrons
   4. electrons

30. Which phrase describes an atom?
   1. a positively charged electron cloud surrounding a positively charged nucleus
   2. a positively charged electron cloud surrounding a negatively charged nucleus
   3. a negatively charged electron cloud surrounding a positively charged nucleus
   4. a negatively charged electron cloud surrounding a negatively charged nucleus
31. Which total mass is the **smallest**?
   1. the mass of 2 electrons
   2. the mass of 2 neutrons
   3. the mass of 1 electron plus the mass of 1 proton
   4. the mass of 1 neutron plus the mass of 1 electron

32. In an atom of argon-40, the number of protons
   1. equals the number of electrons
   2. equals the number of neutrons
   3. is less than the number of electrons
   4. is greater than the number of electrons

33. Which subatomic particles are located in the nucleus of an He-4 atom?
   1. electrons and neutrons
   2. electrons and protons
   3. neutrons and protons
   4. neutrons, protons, and electrons

34. In the late 1800s, experiments using cathode ray tubes led to the discovery of the
   1. electron
   2. neutron
   3. positron
   4. proton

35. The greatest composition by mass in an atom of $^{17}_{8}$O is due to the total mass of its
   1. electrons
   2. neutrons
   3. positrons
   4. protons

36. Describe the electrons in an atom of carbon in the ground state. Your response must include:
   - the charge of an electron [1]
   - the location of electrons based on the wave-mechanical model [1]
   - the total number of electrons in a carbon atom [1]

37. One electron is removed from both an Na atom and a K atom, producing two ions. Using principles of atomic structure, explain why the Na ion is much smaller than the K ion. Discuss both ions in your answer. [2]

38. Base your answer on the information below.

In 1897, J. J. Thomson demonstrated in an experiment that cathode rays were deflected by an electric field. This suggested that cathode rays were composed of negatively charged particles found in all atoms. Thomson concluded that the atom was a positively charged sphere of almost uniform density in which negatively charged particles were embedded. The total negative charge in the atom was balanced by the positive charge, making the atom electrically neutral.

In the early 1900s, Ernest Rutherford bombarded a very thin sheet of gold foil with alpha particles. After interpreting the results of the gold foil experiment, Rutherford proposed a more sophisticated model of the atom.

State one aspect of the modern model of the atom that agrees with a conclusion made by Thomson.
39. Base your answer on the information below.
An atom in an excited state has an electron configuration of 2-7-2.
Explain, in terms of subatomic particles, why this excited atom is electrically neutral. [1]

**Gold foil experiment:**

1. In Rutherford's gold foil experiments, some alpha particles were deflected from their original paths but most passed through the foil with no deflection. Which statement about gold atoms is supported by these experimental observations?
   - 1. Gold atoms consist mostly of empty space.
   - 2. Gold atoms are similar to alpha particles.
   - 3. Alpha particles and gold nuclei have opposite charges.
   - 4. Alpha particles are more dense than gold atoms.

2. Which conclusion was a direct result of the gold foil experiment?
   - 1. An atom is mostly empty space with a dense, positively charged nucleus.
   - 2. An atom is composed of at least three types of subatomic particles.
   - 3. An electron has a positive charge and is located inside the nucleus.
   - 4. An electron has properties of both waves and particles.

3. The gold foil experiment led to the conclusion that each atom in the foil was composed mostly of empty space because most alpha particles directed at the foil
   - 1. passed through the foil
   - 2. remained trapped in the foil
   - 3. were deflected by the nuclei in gold atoms
   - 4. were deflected by the electrons in gold atoms

4. What was concluded about the structure of the atom as the result of the gold foil experiment?
   - 1. A positively charged nucleus is surrounded by positively charged particles.
   - 2. A positively charged nucleus is surrounded by mostly empty space.
   - 3. A negatively charged nucleus is surrounded by positively charged particles.
   - 4. A negatively charged nucleus is surrounded by mostly empty space.

5. Base your answer on the information below.
In the gold foil experiment, a thin sheet of gold was bombarded with alpha particles. Almost all the alpha particles passed straight through the foil. Only a few alpha particles were deflected from their original paths.  
State one conclusion about atomic structure based on the observation that almost all alpha particles passed straight through the foil. [1]

6. Base your answer on the information below.
In the gold foil experiment, a thin sheet of gold was bombarded with alpha particles. Almost all the alpha particles passed straight through the foil. Only a few alpha particles were deflected from their original paths.  
Explain, in terms of charged particles, why some of the alpha particles were deflected. [1]
7. John Dalton was an English scientist who proposed that atoms were hard, indivisible spheres. In the modern model, the atom has a different internal structure.
   a. Identify one experiment that led scientists to develop the modern model of the atom. [1]
   b. Describe this experiment. [1]
   c. State one conclusion about the internal structure of the atom, based on this experiment. [1]

8. In the early 1900s, experiments were conducted to determine the structure of the atom. One of these experiments involved bombarding gold foil with alpha particles. Most alpha particles passed directly through the foil. Some, however, were deflected at various angles. Based on this alpha particle experiment, state two conclusions that were made concerning the structure of an atom.

9. Base your answer on the information below.
   In 1897, J. J. Thomson demonstrated in an experiment that cathode rays were deflected by an electric field. This suggested that cathode rays were composed of negatively charged particles found in all atoms. Thomson concluded that the atom was a positively charged sphere of almost uniform density in which negatively charged particles were embedded. The total negative charge in the atom was balanced by the positive charge, making the atom electrically neutral.

   In the early 1900s, Ernest Rutherford bombarded a very thin sheet of gold foil with alpha particles. After interpreting the results of the gold foil experiment, Rutherford proposed a more sophisticated model of the atom. State one conclusion from Rutherford's experiment that contradicts one conclusion made by Thomson. [1]

Isotopes, mass number and neutrons:

1. Which particles are isotopes of each other?
   (1) \( \frac{1}{1}X \) and \( \frac{3}{1}X \)
   (2) \( \frac{2}{1}X \) and \( \frac{3}{2}X \)
   (3) \( \frac{2}{1}X \) and \( \frac{4}{2}X \)
   (4) \( \frac{3}{1}X \) and \( \frac{3}{2}X \)

2. Atoms of different isotopes of the same element differ in their total number of
   1. electrons
   2. neutrons
   3. protons
   4. valence electrons

3. The total number of protons, electrons, and neutrons in each of four different atoms are shown in the accompanying table. Which two atoms are isotopes of the same element?

<table>
<thead>
<tr>
<th>Atom</th>
<th>Total Number of Protons</th>
<th>Total Number of Electrons</th>
<th>Total Number of Neutrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>6</td>
<td>6</td>
<td>7</td>
</tr>
<tr>
<td>D</td>
<td>6</td>
<td>6</td>
<td>8</td>
</tr>
<tr>
<td>X</td>
<td>7</td>
<td>7</td>
<td>8</td>
</tr>
<tr>
<td>Z</td>
<td>8</td>
<td>8</td>
<td>9</td>
</tr>
</tbody>
</table>

1. A and D
2. A and Z
3. X and D
4. X and Z
4. The isotopes K-37 and K-42 have the same
   1. decay mode 3. mass number for their atoms
   2. bright-line spectrum 4. total number of neutrons in their atoms

5. The accompanying table shows the number of subatomic particles in atom X and in atom Z.

Atom X and atom Z are isotopes of the element

<table>
<thead>
<tr>
<th>Subatomic Particles in Two Atoms</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atom</td>
</tr>
<tr>
<td>--------</td>
</tr>
<tr>
<td>X</td>
</tr>
<tr>
<td>Z</td>
</tr>
</tbody>
</table>

1. aluminum 3. magnesium
2. carbon 4. nitrogen

6. The accepted values for the atomic mass and percent natural abundance of each naturally occurring isotope of silicon are given in the accompanying data table.

Determine the total number of neutrons in an atom of Si-29. [1]

<table>
<thead>
<tr>
<th>Naturally Occurring Isotopes of Silicon</th>
</tr>
</thead>
<tbody>
<tr>
<td>Isotope</td>
</tr>
<tr>
<td>---------</td>
</tr>
<tr>
<td>Si-28</td>
</tr>
<tr>
<td>Si-29</td>
</tr>
<tr>
<td>Si-30</td>
</tr>
</tbody>
</table>

7. Explain, in terms of protons and neutrons, why U-235 and U-238 are different isotopes of uranium.

8. What is the total number of neutrons in an atom of
   \(^{7}\text{Li}\)?
   1. 7
   2. 10

9. All the isotopes of a given atom have
   1. the same mass number and the same atomic number
   2. the same mass number but different atomic numbers
   3. different mass numbers but the same atomic number
   4. different mass numbers and different atomic numbers

10. Atoms of the same element that have different numbers of neutrons are classified as
    1. charged atoms
    2. charged nuclei
    3. isomers
    4. isotopes
11. The number of neutrons in the nucleus of an atom can be determined by
   1. adding the atomic number to the mass number
   2. subtracting the atomic number from the mass number
   3. adding the mass number to the atomic mass
   4. subtracting the mass number from the atomic number

12. What is the total number of neutrons in an atom of $^{57}_{26}$Fe?
   1. 26  
   2. 31  
   3. 57  
   4. 83

13. What is the mass number of a carbon atom that contains six protons, eight neutrons, and six electrons?
   1. 6  
   2. 8  
   3. 14  
   4. 20

14. An atom of potassium-37 and an atom of potassium-42 differ in their total number of
   1. electrons  
   2. neutrons  
   3. protons  
   4. positrons

15. What is the mass number of an atom that has six protons, six electrons, and eight neutrons?
   1. 6  
   2. 12  
   3. 14  
   4. 20

16. What is the mass number of $^{19}_{9}$F?
   1. 9  
   2. 10  
   3. 19  
   4. 28

17. In terms of atomic particles, state one difference between these three isotopes of neon. [1]

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Atomic Mass (atomic mass units)</th>
<th>Percent Natural Abundance</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{20}$Ne</td>
<td>19.99</td>
<td>90.9%</td>
</tr>
<tr>
<td>$^{21}$Ne</td>
<td>20.99</td>
<td>0.3%</td>
</tr>
<tr>
<td>$^{22}$Ne</td>
<td>21.99</td>
<td>8.8%</td>
</tr>
</tbody>
</table>

18. An atom has an atomic number of 9, a mass number of 19, and an electron configuration of 2-6-1. What is the total number of neutrons in this atom?

19. In living organisms, the ratio of the naturally occurring isotopes of carbon, C-12 to C-13 to C-14, is fairly consistent. When an organism such as a woolly mammoth died, it stopped taking in carbon, and the amount of C-14 present in the mammoth began to decrease. For example, one fossil of a woolly mammoth is found to have 1/32 of the amount of C-14 found in a living organism. State, in terms of subatomic particles, how an atom of C-13 is different from an atom of C-12. [1]
20. Base your answer on the accompanying table.
State, in terms of the number of subatomic particles, one similarity and one difference between the atoms of these isotopes of sulfur. [1]

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Atomic Mass (atomic mass units, u)</th>
<th>Natural Abundance (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>^32S</td>
<td>31.97</td>
<td>94.93</td>
</tr>
<tr>
<td>^33S</td>
<td>32.97</td>
<td>0.76</td>
</tr>
<tr>
<td>^34S</td>
<td>33.97</td>
<td>4.29</td>
</tr>
<tr>
<td>^36S</td>
<td>35.97</td>
<td>0.02</td>
</tr>
</tbody>
</table>

21. Base your answers to the question on the information below.
Two isotopes of potassium are K-37 and K-42.
What is the total number of neutrons in the nucleus of a K-37 atom?

22. Base your answers to the question on the information below.
Two isotopes of potassium are K-37 and K-42.
Explain, in terms of subatomic particles, why K-37 and K-42 are isotopes of potassium.

23. The nucleus of an atom of K-42 contains
   1. 19 protons and 23 neutrons
   2. 19 protons and 42 neutrons
   3. 20 protons and 19 neutrons
   4. 23 protons and 19 neutrons

24. Which notation represents an atom of sodium with an atomic number of 11 and a mass number of 24?
   1. \(^{24}_{11}\)Na
   2. \(^{23}_{11}\)Na
   3. \(^{13}_{11}\)Na
   4. \(^{35}_{11}\)Na

Calculating average atomic mass:

1. The atomic mass of an element is calculated using the
   1. atomic number and the ratios of its naturally occurring isotopes
   2. atomic number and the half-lives of each of its isotopes
   3. masses and the ratios of its naturally occurring isotopes
   4. masses and the half-lives of each of its isotopes

2. The atomic mass of an element is the weighted average of the
   1. number of protons in the isotopes of that element
   2. number of neutrons in the isotopes of that element
   3. atomic numbers of the naturally occurring isotopes of that element
   4. atomic masses of the naturally occurring isotopes of that element
3. A 100.00-gram sample of naturally occurring boron contains 19.78 grams of boron-10 (atomic mass = 10.01 atomic mass units) and 80.22 grams of boron-11 (atomic mass = 11.01 atomic mass units). Which numerical setup can be used to determine the atomic mass of naturally occurring boron?

1. $(0.1978)(10.01) + (0.8022)(11.01)$
2. $(0.8022)(10.01) + (0.1978)(11.01)$
3. $[(0.1978)(10.01)]/[(0.8022)(11.01)]$
4. $[(0.8022)(10.01)]/[(0.1978)(11.01)]$

4. What information is necessary to determine the atomic mass of the element chlorine?

1. the atomic mass of each artificially produced isotope of chlorine, only
2. the relative abundance of each naturally occurring isotope of chlorine, only
3. the atomic mass and the relative abundance of each naturally occurring isotope of chlorine
4. the atomic mass and the relative abundance of each naturally occurring and artificially produced isotope of chlorine

5. The atomic mass of titanium is 47.88 atomic mass units. This atomic mass represents the

1. total mass of all the protons and neutrons in an atom of Ti
2. total mass of all the protons, neutrons, and electrons in an atom of Ti
3. weighted average mass of the most abundant isotope of Ti
4. weighted average mass of all the naturally occurring isotopes of Ti

6. Base your answer on the information below.
The accepted values for the atomic mass and percent natural abundance of each naturally occurring isotope of silicon are given in the accompanying data table.
In the space provided, or, if taken online, on a separate piece of paper, show a correct numerical setup for calculating the atomic mass of Si. [1]

<table>
<thead>
<tr>
<th>Naturally Occurring Isotopes of Silicon</th>
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</tr>
<tr>
<td>Si-29</td>
</tr>
<tr>
<td>Si-30</td>
</tr>
</tbody>
</table>

7. Base your answer on the data table, which shows three isotopes of neon.
Based on the atomic masses and the natural abundances shown in the data table, on a separate piece of paper, show a correct numerical setup for calculating the average atomic mass of neon. [1]
8. Base your answer on the data table, which shows three isotopes of neon. Based on natural abundances, the average atomic mass of neon is closest to which whole number? [1]

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<td>0.3%</td>
</tr>
<tr>
<td>$^{22}$Ne</td>
<td>21.99</td>
<td>8.8%</td>
</tr>
</tbody>
</table>

9. Base your answer on the accompanying table. In the space provided or on a separate piece of paper, show a correct numerical setup for calculating the atomic mass of sulfur. [1]

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Atomic Mass (atomic mass units, u)</th>
<th>Natural Abundance (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{32}$S</td>
<td>31.97</td>
<td>94.93</td>
</tr>
<tr>
<td>$^{33}$S</td>
<td>32.97</td>
<td>0.76</td>
</tr>
<tr>
<td>$^{34}$S</td>
<td>33.97</td>
<td>4.29</td>
</tr>
<tr>
<td>$^{36}$S</td>
<td>35.97</td>
<td>0.02</td>
</tr>
</tbody>
</table>

**Ground state and excited state electron configuration:**

1. Which ion has the same electron configuration as an atom of He?
   
   1. $\text{H}^-$
   
   2. $\text{O}^2-$
   
   3. $\text{Na}^+$
   
   4. $\text{Ca}^{2+}$

2. Which electron configuration could represent a strontium atom in an excited state?
   
   1. 2-8-18-7-1
   
   2. 2-8-18-7-3
   
   3. 2-8-18-8-1
   
   4. 2-8-18-8-2

3. Compared to an electron in the first electron shell of an atom, an electron in the third shell of the same atom has
   
   1. less mass
   
   2. less energy
   
   3. more mass
   
   4. more energy

4. Which electron configuration represents an atom in an excited state?
   
   1. 2-7
   
   2. 2-6-2
   
   3. 2-8-1
   
   4. 2-8-8-2
5. The light emitted from a flame is produced when electrons in an excited state
   1. absorb energy as they move to lower energy states
   2. absorb energy as they move to higher energy states
   3. release energy as they move to lower energy states
   4. release energy as they move to higher energy states

6. How do the energy and the most probable location of an electron in the third shell of an atom compare to the energy and the most probable location of an electron in the first shell of the same atom?
   1. In the third shell, an electron has more energy and is closer to the nucleus.
   2. In the third shell, an electron has more energy and is farther from the nucleus.
   3. In the third shell, an electron has less energy and is closer to the nucleus.
   4. In the third shell, an electron has less energy and is farther from the nucleus.

7. What is the total number of protons in an atom with the electron configuration 2-8-18-32-18-1?
   1. 69
   2. 79
   3. 118
   4. 197

8. An electron in a sodium atom moves from the third shell to the fourth shell. This change is a result of the atom
   1. absorbing energy
   2. releasing energy
   3. gaining an electron
   4. losing an electron

9. Which electron configuration represents an excited state for a potassium atom?
   1. 2-8-7-1
   2. 2-8-7-2
   3. 2-8-8-1
   4. 2-8-8-2

10. Given the bright-line spectra of three elements and the spectrum of a mixture formed from at least two of these elements:

    Which elements are present in this mixture?

    ![Bright-Line Spectra](image)

    1. E and D, only
    2. E and G, only
    3. D and G, only
    4. D, E, and G
11. Base your answer on the information below.
In a laboratory, a glass tube is filled with hydrogen gas at a very low pressure. When a scientist applies high voltage between metal electrodes in the tube, light is emitted. The scientist analyzes the light with a spectroscope and observes four distinct spectral lines. The accompanying table gives the color, frequency, and energy for each of the four spectral lines. The unit for frequency is hertz, Hz. Explain, in terms of subatomic particles and energy states, why light is emitted by the hydrogen gas.

<table>
<thead>
<tr>
<th>Color</th>
<th>Frequency (×10^14 Hz)</th>
<th>Energy (×10^-19 J)</th>
</tr>
</thead>
<tbody>
<tr>
<td>red</td>
<td>4.6</td>
<td>3.0</td>
</tr>
<tr>
<td>blue green</td>
<td>6.2</td>
<td>4.1</td>
</tr>
<tr>
<td>blue</td>
<td>6.9</td>
<td>4.6</td>
</tr>
<tr>
<td>violet</td>
<td>7.3</td>
<td>4.8</td>
</tr>
</tbody>
</table>

12. On a separate piece of paper, write an electron configuration for a silicon atom in an excited state.

13. Base your answer on the accompanying information.
The accompanying bright-line spectra for three elements and a mixture of elements are shown.
   a Identify all the elements in the mixture. [1]
   b Explain, in terms of both electrons and energy, how the bright-line spectrum of an element is produced. [1]

14. Write an electron configuration for an atom of aluminum-27 in an excited state. [1]

15. Base your answer on the diagram shown, which represents an atom of magnesium-26 in the ground state.
   Copy diagram B on a separate piece of paper and write an appropriate number of electrons in each shell to represent a Mg-26 atom in an excited state. Your answer may include additional shells.
16. Base your answer on the information below.
An atom has an atomic number of 9, a mass number of 19, and an electron configuration of 2-6-1.
Explain why the number of electrons in the second and third shells shows that this atom is in an excited state.

17. Base your answer on the information below.
A thiol is very similar to an alcohol, but a thiol has a sulfur atom instead of an oxygen atom in the functional group. One of the compounds in a skunk's spray is 2-butene-1-thiol. The formula of this compound is shown in the accompanying diagram.
Explain, in terms of electron configuration, why oxygen atoms and sulfur atoms form compounds with similar molecular structures.

18. Base your answer on the information below.
A glass tube is filled with hydrogen gas at low pressure. An electric current is passed through the gas, causing it to emit light. This light is passed through a prism to separate the light into the bright, colored lines of hydrogen's visible spectrum. Each colored line corresponds to a particular wavelength of light. One of hydrogen's spectral lines is red light with a wavelength of 656 nanometers.
Tubes filled with other gases produce different bright-line spectra that are characteristic of each kind of gas. These spectra have been observed and recorded.

a Explain, in terms of electron energy states and energy changes, how hydrogen's bright-line spectrum is produced.

b Explain how the elements present on the surface of a star can be identified using bright-line spectra.
19. Base your answer on the information below:
The Balmer series refers to the visible bright lines in the spectrum produced by hydrogen atoms. The color and wavelength of each line in this series are given in the accompanying table. Explain, in terms of both subatomic particles and energy states, how the Balmer series is produced. [1]

20. Base your answer on the information below.
An atom in an excited state has an electron configuration of 2-7-2. Write the electron configuration of this atom in the ground state. [1]

Valence electrons and Lewis structures:

1. What is represented by the dots in a Lewis electron-dot diagram of an atom of an element in Period 2 of the Periodic Table?
   1. the number of neutrons in the atom
   2. the number of protons in the atom
   3. the number of valence electrons in the atom
   4. the total number of electrons in the atom

2. An atom in the ground state contains a total of 5 electrons, 5 protons, and 5 neutrons. Which Lewis electron-dot diagram represents this atom?

   ![](image)

   (1) (2) (3) (4)

3. Which Lewis electron-dot diagram represents an atom in the ground state for a Group 13 element?

   ![](image)

   (1) (2) (3) (4)

   1. (1)
   2. (2)
   3. (3)
   4. (4)

4. What is the total number of valence electrons in a calcium atom in the ground state?

   1. 8
   2. 2
   3. 18
   4. 20

5. Base your answers to the question on the information below.
Two isotopes of potassium are K-37 and K-42. How many valence electrons are in an atom of K-42 in the ground state?

6. Base your answer on the accompanying information.
The accompanying bright-line spectra for three elements and a mixture of elements are shown. State the total number of valence electrons in a cadmium atom in the ground state. [1]
7. In the space provided or on a separate sheet of paper, draw a Lewis electron-dot diagram for a sulfur atom in the ground state. [1]

8. In the box shown, draw the electron-dot (Lewis) structure of an atom of chlorine. [1]

9. In the box shown, draw the electron-dot (Lewis) structure of an atom of calcium. [1]

10. Base your answer on the electron configuration table. What is the total number of valence electrons in an atom of electron configuration \( X \)? [1]

<table>
<thead>
<tr>
<th>Element</th>
<th>Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>( X )</td>
<td>2–8–8–2</td>
</tr>
<tr>
<td>( Y )</td>
<td>2–8–7–3</td>
</tr>
<tr>
<td>( Z )</td>
<td>2–8–8</td>
</tr>
</tbody>
</table>
11. Base your answer on the diagram shown, which represents an atom of magnesium-26 in the ground state.
What is the total number of valence electrons in an atom of Mg-26 in the ground state?

![Diagram of Mg-26](image)

12. Base your answer on the accompanying table.
*In the space provided or on a separate piece of paper*, draw a Lewis electron-dot diagram for an atom of sulfur-33. [1]

<table>
<thead>
<tr>
<th>Naturally Occurring Isotopes of Sulfur</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Isotope</strong></td>
</tr>
<tr>
<td>^{32}\text{S}</td>
</tr>
<tr>
<td>^{33}\text{S}</td>
</tr>
<tr>
<td>^{34}\text{S}</td>
</tr>
<tr>
<td>^{36}\text{S}</td>
</tr>
</tbody>
</table>

**Radiation:**

1. The accompanying diagram represents radioactive emanations passing through an electric field. Which type of emanation is represented by the arrow labeled 1?

![Diagram of Radioactive Emanations](image)

1. alpha particle  
2. beta particle  
3. positron  
4. gamma ray

2. Which particle has the least mass?

1. alpha particle  
2. beta particle  
3. neutron  
4. proton

3. Which particle has the least mass?

1. \( ^{4}\text{He} \)  
2. \( ^{1}\text{H} \)  
3. \( ^{1}\text{H} \)  
4. \( ^{1}\text{H} \)
**Periodic Table and Table S:**

1. Which element has the highest electrical conductivity?
   1. Mg  
   2. H  
   3. He  
   4. Cl

2. What is the total number of valence electrons in a fluorine atom in the ground state?
   1. 5  
   2. 2  
   3. 7  
   4. 9

3. The accompanying graph represents the relationship between atomic radii, in picometers, and increasing atomic number for elements in Group 15. Which element is most metallic?

   ![Graph showing atomic radii vs atomic number]

   1. A  
   2. B  
   3. D  
   4. E

4. Which element of Group 17 exists as a solid at 25°C and standard pressure?
   1. fluorine  
   2. chlorine  
   3. bromine  
   4. iodine

5. Which group in the Periodic Table contains elements that are all monatomic gases at STP?
   1. 15  
   2. 16  
   3. 17  
   4. 18

6. Which molecule contains a triple covalent bond between its atoms?
   1. N₂  
   2. O₂  
   3. F₂  
   4. H₂

7. Which element is classified as a noble gas at STP?
   1. hydrogen  
   2. oxygen  
   3. neon  
   4. nitrogen
8. Metallic bonding occurs between atoms of
   1. sulfur       3. fluorine
   2. copper       4. carbon

9. In which shell are the valence electrons of the elements in Period 2 found?
   1. 1           3. 3
   2. 2           4. 4

10. Which of the following Group 15 elements has the greatest metallic character?
    1. nitrogen     3. antimony
    2. phosphorus   4. bismuth

11. Which of the following ions has the smallest radius?
    1. F⁻           3. K⁺
    2. Cl⁻           4. Ca²⁺

12. The atomic number of an atom is always equal to the number of its
    1. protons, only     3. protons plus neutrons
    2. neutrons, only    4. protons plus electrons

13. Which list of elements is arranged in order of increasing atomic radii?
    1. Li, Be, B, C     3. Sc, Ti, V, Cr
    2. Sr, Ca, Mg, Be   4. F, Cl, Br, I

14. What is the charge of the nucleus in an atom of oxygen-17?
    1. 0              3. +8
    2. -2             4. +17

15. Compared to an atom of phosphorus-31, an atom of sulfur-32 contains
    1. one less neutron     3. one more neutron
    2. one less proton      4. one more proton

16. 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     3. Group 16
    2. Group 2     4. Group 17

17. At standard pressure, which element has a melting point higher than standard temperature?
    1. F₂           3. Fe
    2. Br₂           4. Hg
18. A sample composed only of atoms having the same atomic number is classified as
   1. a compound
   2. a solution
   3. an element
   4. an isomer

19. 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?

   ![Diagram of circles representing atom sizes]

   (1) (2) (3) (4)

20. An atom in the ground state has seven valence electrons. This atom could be an atom of which
    element?
   1. calcium
   2. fluorine
   3. oxygen
   4. sodium

21. Which statement identifies the element arsenic?
   1. Arsenic has an atomic number of 33.
   2. Arsenic has a melting point of 84 K.
   3. An atom of arsenic in the ground state has eight valence electrons.
   4. An atom of arsenic in the ground state has a radius of 146 pm.

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

23. Which element is a metalloid?
   1. Al
   2. Ar
   3. As
   4. Au

24. The compound XCl is classified as ionic if X represents the element
   1. H
   2. I
   3. Rb
   4. Br

25. Which Group 14 element is classified as a metal?
   1. carbon
   2. germanium
   3. silicon
   4. tin
26. Which element forms a compound with chlorine with the general formula $M\text{Cl}$?
   1. Rb  
   2. Ra  
   3. Re  
   4. Rn

27. An atom of argon in the ground state tends *not* to bond with an atom of a different element because the argon atom has
   1. more protons than neutrons  
   2. more neutrons than protons  
   3. a total of two valence electrons  
   4. a total of eight valence electrons

28. A sample of matter must be copper if
   1. each atom in the sample has 29 protons  
   2. atoms in the sample react with oxygen  
   3. the sample melts at 1768 K  
   4. the sample can conduct electricity

29. The elements on the Periodic Table are arranged in order of increasing
   1. atomic number  
   2. mass number  
   3. number of isotopes  
   4. number of moles

30. Which element has the highest melting point?
   1. tantalum  
   2. rhenium  
   3. osmium  
   4. hafnium

31. Compared to the atoms of nonmetals in Period 3, the atoms of metals in Period 3 have
   1. fewer valence electrons  
   2. more valence electrons  
   3. fewer electron shells  
   4. more electron shells

32. Which atom in the ground state requires the *least* amount of energy to remove its valence electron?
   1. lithium atom  
   2. potassium atom  
   3. rubidium atom  
   4. sodium atom

33. What is the total number of protons in the nucleus of an atom of potassium-42?
   1. 15  
   2. 19  
   3. 39  
   4. 42

34. In which group of the Periodic Table do most of the elements exhibit both positive and negative oxidation states?
   1. 17  
   2. 2  
   3. 12  
   4. 7
35. What determines the order of placement of the elements on the modern Periodic Table?
   1. atomic number
   2. atomic mass
   3. the number of neutrons, only
   4. the number of neutrons and protons

36. Which statement best describes the shape and volume of an aluminum cylinder at STP?
   1. It has a definite shape and a definite volume.
   2. It has a definite shape and no definite volume.
   3. It has no definite shape and a definite volume.
   4. It has no definite shape and no definite volume.

37. Which diagram represents the nucleus of an atom of $^{27}_{13}$Al?

```
(1) 14 n 27 p
   14 n 13 p

(2) 27 n 13 p
(3) 40 n 13 p
```

38. Which list of elements from Group 2 on the Periodic Table is arranged in order of increasing atomic radius?
   1. Be, Mg, Ca
   2. Ca, Mg, Be
   3. Ba, Ra, Sr
   4. Sr, Ra, Ba

39. Which element is a solid at STP and a good conductor of electricity?
   1. iodine
   2. mercury
   3. nickel
   4. sulfur

40. Which statement explains why sulfur is classified as a Group 16 element?
   1. A sulfur atom has 6 valence electrons.
   2. A sulfur atom has 16 neutrons.
   3. Sulfur is a yellow solid at STP.
   4. Sulfur reacts with most metals.

41. Which group on the Periodic Table of the Elements contains elements that react with oxygen to form compounds with the general formula $X_2O$?
   1. Group 1
   2. Group 2
   3. Group 14
   4. Group 18

42. Elements on the modern Periodic Table are arranged in order of increasing
   1. atomic mass
   2. atomic number
   3. number of neutrons
   4. number of valence electrons
43. Chlorine-37 can be represented as

(1) $^{17}_{35}$Cl  
(2) $^{35}_{17}$Cl  
(3) $^{37}_{17}$Cl  
(4) $^{35}_{20}$Cl

44. Which element is a metal that is in the liquid phase at STP?

1. bromine  
2. cobalt  
3. hydrogen  
4. mercury

45. In the ground state, each atom of an element has two valence electrons. This element has a lower first ionization energy than calcium. Where is the element located on the Periodic Table?

1. Group 1, Period 4  
2. Group 2, Period 5  
3. Group 2, Period 3  
4. Group 3, Period 4

46. Which balanced equation represents a chemical change?

(1) $\text{H}_2\text{O}(\ell) + \text{energy} \rightarrow \text{H}_2\text{O}(g)$  
(2) $2\text{H}_2\text{O}(\ell) + \text{energy} \rightarrow 2\text{H}_2(g) + \text{O}_2(g)$  
(3) $\text{H}_2\text{O}(\ell) \rightarrow \text{H}_2\text{O}(s) + \text{energy}$  
(4) $\text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(\ell) + \text{energy}$

47. An atom of aluminum in the ground state and an atom of gallium in the ground state have the same

1. mass  
2. electronegativity  
3. total number of protons  
4. total number of valence electrons

48. Which element is a liquid at 758 K and standard pressure?

1. gold  
2. silver  
3. platinum  
4. thallium

49. An atom of which element has the largest atomic radius?

1. Fe  
2. Mg  
3. Si  
4. Zn

50. Which element requires the least amount of energy to remove the most loosely held electron from a gaseous atom in the ground state?

1. bromine  
2. calcium  
3. sodium  
4. silver

51. An atom of helium-4 differs from an atom of lithium-7 in that the atom of helium-4 has

1. one more proton  
2. one more neutron  
3. two less protons  
4. two less neutrons
52. Base your answer on the information below and the accompanying table. The table shown lists physical and chemical properties of six elements at standard pressure that correspond to known elements on the Periodic Table. The elements are identified by the code letters, D, E, G, J, L, and Q.

<table>
<thead>
<tr>
<th>Properties of Six Elements at Standard Pressure</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Element D</strong></td>
</tr>
<tr>
<td>Density 0.00018 g/cm³</td>
</tr>
<tr>
<td>Melting point –272°C</td>
</tr>
<tr>
<td>Boiling point –269°C</td>
</tr>
<tr>
<td>Oxide formula (none)</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th><strong>Element J</strong></th>
<th><strong>Element L</strong></th>
<th><strong>Element Q</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>Density 0.0013 g/cm³</td>
<td>Density 0.86 g/cm³</td>
<td>Density 0.97 g/cm³</td>
</tr>
<tr>
<td>Melting point –210°C</td>
<td>Melting point 64°C</td>
<td>Melting point 98°C</td>
</tr>
<tr>
<td>Boiling point –196°C</td>
<td>Boiling point 774°C</td>
<td>Boiling point 863°C</td>
</tr>
<tr>
<td>Oxide formula J₂O₅</td>
<td>Oxide formula L₂O</td>
<td>Oxide formula Q₂O</td>
</tr>
</tbody>
</table>

a) What is the total number of elements in the "Properties of Six Elements at Standard Pressure" table that are solids at STP? [1]
b) An atom of element G is in the ground state. What is the total number of valence electrons in this atom? [1]
c) Identify, by code letter, the element that is a noble gas in the "Properties of Six Elements at Standard Pressure" table. [1]

53. Base your answer on the information below. Ozone gas, O₃, can be used to kill adult insects in storage bins for grain without damaging the grain. The ozone is produced from oxygen gas, O₂, in portable ozone generators located near the storage bins. The concentrations of ozone used are so low that they do not cause any environmental damage. This use of ozone is safer and more environmentally friendly than a method that used bromomethane, CH₃Br. However, bromomethane was more effective than ozone because CH₃Br killed immature insects as well as adult insects.

*Adapted From: *The Sunday Gazette* (Schenectady, NY) 3/9/03

Based on the information in the passage, state one advantage of using ozone instead of bromomethane for insect control in grain storage bins. [1]

54. Base your answer on the information below. Arsenic is often obtained by heating the ore arsenopyrite, FeAsS. The decomposition of FeAsS is represented by the balanced equation shown.

\[
\text{FeAsS(s)} \xrightarrow{\text{heat}} \text{FeS(s)} + \text{As(g)}
\]

In the solid phase, arsenic occurs in two forms. One form, yellow arsenic, has a density of 1.97 g/cm³ at STP. The other form, gray arsenic, has a density of 5.78 g/cm³ at STP. When arsenic is heated rapidly in air, arsenic(III) oxide is formed. Although arsenic is toxic, it is needed by the human body in very small amounts. The body of a healthy human adult contains approximately 5 milligrams of arsenic. Explain, in terms of the arrangement of atoms, why the two forms of arsenic have different densities at STP. [1]
55. In the 19th century, Dmitri Mendeleev predicted the existence of a then unknown element $X$ with a mass of 68. He also predicted that an oxide of $X$ would have the formula $X_2O_3$. On the modern Periodic Table, what is the group number and period number of element $X$? [1]

56. Base your answer on the accompanying table. Explain, in terms of atomic structure, why cesium has a lower first ionization energy than rubidium.

**First Ionization Energy of Selected Elements**

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic Number</th>
<th>First Ionization Energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>lithium</td>
<td>3</td>
<td>520</td>
</tr>
<tr>
<td>sodium</td>
<td>11</td>
<td>496</td>
</tr>
<tr>
<td>potassium</td>
<td>19</td>
<td>419</td>
</tr>
<tr>
<td>rubidium</td>
<td>37</td>
<td>403</td>
</tr>
<tr>
<td>cesium</td>
<td>55</td>
<td>376</td>
</tr>
</tbody>
</table>

57. Base your answer on the information below. A metal, $M$, was obtained from a compound in a rock sample. Experiments have determined that the element is a member of Group 2 on the Periodic Table of the Elements. What is the phase of element $M$ at STP? [1]

58. Identify the element in Period 3 of the Periodic Table that reacts with oxygen to form an ionic compound represented by the formula $X_2O$. [1]

59. Base your answer on the accompanying information. Identify one element from this table for each type of element: metal, metalloid, and nonmetal. [1]
60. Base your answer on the elements in Group 2 on the Periodic Table. State, in terms of the number of electron shells, why the radius of a strontium atom in the ground state is larger than the radius of a magnesium atom in the ground state. [1]

61. Base your answer on the information below. The radioisotope uranium-238 occurs naturally in Earth's crust. The disintegration of this radioisotope is the first in a series of spontaneous decays. The sixth decay in this series produces the radioisotope radon-222. The decay of radon-222 produces the radioisotope polonium-218 that has a half life of 3.04 minutes. Eventually, the stable isotope lead-206 is produced by the alpha decay of an unstable nuclide. Explain, in terms of electron configuration, why atoms of the radioisotope produced by the sixth decay in the U-238 disintegration series do not readily react to form compounds. [1]

**Trends:**

1. As the atoms in Period 3 of the Periodic Table are considered from left to right, the atoms generally show
   1. an increase in radius and an increase in ionization energy
   2. an increase in radius and a decrease in ionization energy
   3. a decrease in radius and an increase in ionization energy
   4. a decrease in radius and a decrease in ionization energy

2. Which trends are observed as each of the elements within Group 15 on the Periodic Table is considered in order from top to bottom?
   1. Their metallic properties decrease and their atomic radii decrease.
   2. Their metallic properties decrease and their atomic radii increase.
   3. Their metallic properties increase and their atomic radii decrease.
   4. Their metallic properties increase and their atomic radii increase.

3. Which characteristics both generally decrease when the elements in Period 3 on the Periodic Table are considered in order from left to right?
   1. nonmetallic properties and atomic radius
   2. nonmetallic properties and ionization energy
   3. metallic properties and atomic radius
   4. metallic properties and ionization energy

4. As the elements of Group 17 are considered in order of increasing atomic number, there is an increase in
   1. atomic radius
   2. electronegativity
   3. first ionization energy
   4. number of electrons in the first shell

5. Which general trend is demonstrated by the Group 17 elements as they are considered in order from top to bottom on the Periodic Table?
   1. a decrease in atomic radius
   2. a decrease in electronegativity
   3. an increase in first ionization energy
   4. an increase in nonmetallic behavior
6. Base your answer on the information below and the accompanying table. The table shown lists physical and chemical properties of six elements at standard pressure that correspond to known elements on the Periodic Table. The elements are identified by the code letters, D, E, G, J, L, and Q. Letter Z corresponds to an element on the Periodic Table other than the six listed elements. Elements G, Q, L, and Z are in the same group on the Periodic Table, as shown in diagram A. Based on the trend in the melting points for elements G, Q, and L listed in the "Properties of Six Elements at Standard Pressure" table, estimate the melting point of element Z, in degrees Celsius. [1]

### Properties of Six Elements at Standard Pressure

<table>
<thead>
<tr>
<th>Element D</th>
<th>Element E</th>
<th>Element G</th>
<th>Dia. A</th>
</tr>
</thead>
<tbody>
<tr>
<td>Density 0.0018 g/cm³</td>
<td>Density 1.82 g/cm³</td>
<td>Density 0.53 g/cm³</td>
<td></td>
</tr>
<tr>
<td>Melting point –272°C</td>
<td>Melting point 44°C</td>
<td>Melting point 181°C</td>
<td></td>
</tr>
<tr>
<td>Boiling point –269°C</td>
<td>Boiling point 280°C</td>
<td>Boiling point 1347°C</td>
<td></td>
</tr>
<tr>
<td>Oxide formula (none)</td>
<td>Oxide formula E₂O₅</td>
<td>Oxide formula G₂O₅</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Element J</th>
<th>Element L</th>
<th>Element Q</th>
</tr>
</thead>
<tbody>
<tr>
<td>Density 0.0013 g/cm³</td>
<td>Density 0.86 g/cm³</td>
<td>Density 0.97 g/cm³</td>
</tr>
<tr>
<td>Melting point –210°C</td>
<td>Melting point 64°C</td>
<td>Melting point 98°C</td>
</tr>
<tr>
<td>Boiling point –196°C</td>
<td>Boiling point 774°C</td>
<td>Boiling point 893°C</td>
</tr>
<tr>
<td>Oxide formula J₂O₅</td>
<td>Oxide formula L₂O</td>
<td>Oxide formula Q₂O</td>
</tr>
</tbody>
</table>

7. Base your answer on the elements in Group 2 on the Periodic Table. State the general trend in first ionization energy for the elements in Group 2 as these elements are considered in order from top to bottom in the group. [1]

### Most similar properties / same group:

1. Arsenic and silicon are similar in that they both
   1. have the same ionization energy
   2. have the same covalent radius
   3. are transition metals
   4. are metalloids

2. Which element has chemical properties that are most similar to those of calcium?
   1. Co
   2. K
   3. N
   4. Sr

3. Which element is most chemically similar to chlorine?
   1. Ar
   2. F
   3. Fr
   4. S

4. Magnesium and calcium have similar chemical properties because an atom of each element has the same total number of
   1. electron shells
   2. valence electrons
   3. neutrons
   4. protons
5. Magnesium and calcium have similar chemical properties because a magnesium atom and a calcium atom have the same:
   1. atomic number
   2. mass number
   3. total number of electron shells
   4. total number of valence electrons

6. Which of these elements has physical and chemical properties most similar to silicon (Si)?
   1. germanium (Ge)
   2. lead (Pb)
   3. phosphorus (P)
   4. chlorine (Cl)

7. Which two elements have the most similar chemical properties?
   1. Be and Mg
   2. Ca and Br
   3. Cl and Ar
   4. Na and P

8. Based on the Periodic Table, explain why Na and K have similar chemical properties. [1]

9. Explain, in terms of atomic structure, why germanium is chemically similar to silicon.

10. Base your answer on the article below, the Reference Tables for Physical Setting/Chemistry, and your knowledge of chemistry.
    In the 1920s, paint used to inscribe the numbers on watch dials was composed of a luminescent (glow-in-the-dark) mixture. The powdered-paint base was a mixture of radium salts and zinc sulfide. As the paint was mixed, the powdered base became airborne and drifted throughout the workroom causing the contents of the workroom, including the painters' clothes and bodies, to glow in the dark. The paint is luminescent because radiation from the radium salts strikes a scintillator. A scintillator is a material that emits visible light in response to ionizing radiation. In watchdial paint, zinc sulfide acts as the scintillator.
    Radium present in the radium salts decomposes spontaneously, emitting alpha particles. These particles can cause damage to the body when they enter human tissue. Alpha particles are especially harmful to the blood, liver, lungs, and spleen because they can alter genetic information in the cells. Radium can be deposited in the bones because it substitutes for calcium.

   Why does radium substitute for calcium in bones? [1]

11. Base your answer on the reading passage shown and on your knowledge of chemistry.
    **A Glow in the Dark, and Scientific Peril**
    The [Marie and Pierre] Curies set out to study radioactivity in 1898. Their first accomplishment was to show that radioactivity was a property of atoms themselves. Scientifically, that was the most important of their findings, because it helped other researchers refine their understanding of atomic structure.
    More famous was their discovery of polonium and radium. Radium was the most radioactive substance the Curies had encountered. Its radioactivity is due to the large size of the atom, which makes the nucleus unstable and prone to decay, usually to radon and then lead, by emitting particles and energy as it seeks a more stable configuration.
    Marie Curie struggled to purify radium for medical uses, including early radiation treatment for tumors. But radium's bluish glow caught people's fancy, and companies in the United States began mining it and selling it as a novelty: for glow-in-the-dark light pulls, for instance, and bogus cure-all patent medicines that actually killed people.
What makes radium so dangerous is that it forms chemical bonds in the same way as calcium, and the body can mistake it for calcium and absorb it into the bones. Then, it can bombard cells with radiation at close range, which may cause bone tumors or bone-marrow damage that can give rise to anemia or leukemia.


Using information from the Periodic Table, explain why radium forms chemical bonds in the same way as calcium does. [1]

12. Explain, in terms of electron configuration, why selenium and sulfur have similar chemical properties. [1]

13. Base your answer on the information below. Elements with atomic numbers 112 and 114 have been produced and their IUPAC names are pending approval. However, an element that would be put between these two elements on the Periodic Table has not yet been produced. If produced, this element will be identified by the symbol Uut until an IUPAC name is approved. Identify one element that would be chemically similar to Uut. [1]

14. Base your answer on the elements in Group 2 on the Periodic Table. Explain, in terms of atomic structure, why the elements in Group 2 have similar chemical properties.

Properties of elements:

1. Which statement is true about the properties of the elements in any one period of the Periodic Table?
   1. They are determined by the number of neutrons.
   2. They are determined by the number of electrons in the first shell.
   3. They change in a generally systematic manner.
   4. They change in a random, unpredictable manner.

2. What are two properties of most nonmetals?
   1. high ionization energy and poor electrical conductivity
   2. high ionization energy and good electrical conductivity
   3. low ionization energy and poor electrical conductivity
   4. low ionization energy and good electrical conductivity

3. Which is a property of most nonmetallic solids?
   1. high thermal conductivity
   2. high electrical conductivity
   3. brittleness
   4. malleability

4. 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.
5. An example of a physical property of an element is the element’s ability to
   1. react with an acid  
   2. react with oxygen  
   3. form a compound with chlorine  
   4. form an aqueous solution

6. Which statement correctly describes two forms of oxygen, O₂ and O₃?
   1. They have identical molecular structures and identical properties.  
   2. They have identical molecular structures and different properties.  
   3. They have different molecular structures and identical properties.  
   4. They have different molecular structures and different properties.

7. Which two characteristics are associated with metals?
   1. low first ionization energy and low electronegativity  
   2. low first ionization energy and high electronegativity  
   3. high first ionization energy and low electronegativity  
   4. high first ionization energy and high electronegativity

8. At STP, which element is brittle and not a conductor of electricity?
   1. S  
   2. K  
   3. Na  
   4. Ar

9. 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

10. At STP, solid carbon can exist as diamond and graphite. Compared to the molecular structure and chemical properties of diamond, graphite has
   1. a different molecular structure and different properties  
   2. a different molecular structure and the same properties  
   3. the same molecular structure and different properties  
   4. the same molecular structure and the same properties

11. At STP, both diamond and graphite are solids composed of carbon atoms. These solids have
   1. the same crystal structure and the same properties  
   2. the same crystal structure and different properties  
   3. different crystal structures and the same properties  
   4. different crystal structures and different properties

12. Which elements are malleable and good conductors of electricity?
   1. iodine and silver  
   2. iodine and xenon  
   3. tin and silver  
   4. tin and xenon
13. At 298 K, oxygen (O\(_2\)) and ozone (O\(_3\)) have different properties because their
   1. atoms have different atomic numbers    3. molecules have different molecular structures
   2. atoms have different atomic masses    4. molecules have different average kinetic energies

14. Which Period 4 element has the most metallic properties?
   1. As    3. Ge
   2. Br    4. Se

15. Which statement explains why ozone gas, O\(_3\), and oxygen gas, O\(_2\), have different properties?
   1. They are formed from different elements. 3. They have different oxidation numbers.
   2. They have different molecular structures. 4. They have different electronegativities.

16. Which element has both metallic and nonmetallic properties?
   1. Rb    3. Si
   2. Rn    4. Sr

17. The carbon atoms in graphite and the carbon atoms in diamond have different
   1. atomic numbers    3. electronegativities
   2. atomic masses    4. structural arrangements

18. Which statement describes a chemical property of the element magnesium?
   1. Magnesium is malleable. 3. Magnesium reacts with an acid.
   2. Magnesium conducts electricity. 4. Magnesium has a high boiling point.

19. How do the atomic radius and metallic properties of sodium compare to the atomic radius and
    metallic properties of phosphorus?
   1. Sodium has a larger atomic radius and is more metallic. 3. Sodium has a smaller atomic radius and is
     more metallic.
   2. Sodium has a larger atomic radius and is less metallic. 4. Sodium has a smaller atomic radius and is
     less metallic.

20. Which statement describes oxygen gas, O\(_2\)(g), and ozone gas, O\(_3\)(g)?
   1. They have different molecular structures, only. 3. They have different molecular structures and different properties.
   2. They have different properties, only. 4. They have the same molecular structure and the same properties.

21. Which statement describes a chemical property of bromine?
   1. Bromine is soluble in water. 3. Bromine combines with aluminum to produce AlBr\(_3\).
   2. Bromine has a reddish-brown color. 4. Bromine changes from a liquid to a gas at 332 K and 1 atm.
22. A sample of an element is malleable and can conduct electricity. This element could be
   1. H  3. S
   2. He 4. Sn

23. At STP, which element is solid, brittle, and a poor conductor of electricity?
   1. Al 3. Ne
   2. K  4. S

24. Which element is a brittle, nonconducting solid at 25°C?
   1. Br 3. Al
   2. S 4. Bi

25. Two sources of copper are cuprite, which has the IUPAC name copper(I) oxide, and malachite, which has the formula Cu₂CO₃(OH)₂. Copper is used in home wiring and electric motors because it has good electrical conductivity. Other uses of copper not related to its electrical conductivity include coins, plumbing, roofing, and cooking pans. Aluminum is also used for cooking pans. At room temperature, the electrical conductivity of a copper wire is 1.6 times greater than an aluminum wire with the same length and cross-sectional area. At room temperature, the heat conductivity of copper is 1.8 times greater than the heat conductivity of aluminum. At STP, the density of copper is 3.3 times greater than the density of aluminum.
   a) Identify one physical property of copper that makes it a good choice for uses that are not related to electrical conductivity. [1]
   b) Identify one physical property of aluminum that could make it a better choice than copper for a cooking pan. [1]

26. Base your answer to the question on the information below.
   Element X is a solid metal that reacts with chlorine to form a water-soluble binary compound.
   State one physical property of element X that makes it a good material for making pots and pans.

27. Base your answer on the information below.
   Carbon and oxygen are examples of elements that exist in more than one form in the same phase.
   Graphite and diamond are two crystalline arrangements for carbon. The crystal structure of graphite is organized in layers. The bonds between carbon atoms within each layer of graphite are strong. The bonds between carbon atoms that connect different layers of graphite are weak because the shared electrons in these bonds are loosely held by the carbon atoms. The crystal structure of diamond is a strong network of atoms in which all the shared electrons are strongly held by the carbon atoms. Graphite is an electrical conductor, but diamond is not. At 25°C, graphite has a density of 2.2 g/cm³ and diamond has a density of 3.51 g/cm³.
   The element oxygen can exist as diatomic molecules, O₂, and as ozone, O₃. At standard pressure the boiling point of ozone is 161 K.
   Explain, in terms of electrons, why graphite is an electrical conductor and diamond is not. Your response must include information about both graphite and diamond. [1]
**Chemical Bonding:**

1. Given the balanced equation representing a reaction: \( \text{Cu} + \text{S} \rightarrow \text{CuS} + \text{energy} \)
Which statement explains why the energy term is written to the right of the arrow?

   1. The compound CuS is composed of two metals.
   2. The compound CuS is composed of two nonmetals.
   3. Energy is absorbed as the bonds in CuS form.
   4. Energy is released as the bonds in CuS form.

2. Base your answer on the information below.
Naphthalene, a nonpolar substance that sublimes at room temperature, can be used to protect wool clothing from being eaten by moths.
Explain why naphthalene is not expected to dissolve in water. [1]

3. Which Lewis electron-dot diagram represents calcium oxide?

4. Which compound has molecules that form the strongest hydrogen bonds?

   1. HI
   2. HBr
   3. HF
   4. HCl

5. Hexane \((\text{C}_6\text{H}_{14})\) and water do not form a solution. Which statement explains this phenomenon?

   1. Hexane is polar and water is nonpolar.
   2. Hexane is nonpolar and water is polar.
   3. Hexane is nonpolar and water is ionic.
   4. Hexane is nonpolar and water is ionic.

6. Based on bond type, which compound has the highest melting point?

   1. \(\text{CH}_3\text{OH}\)
   2. \(\text{C}_6\text{H}_{14}\)
   3. \(\text{CaCl}_2\)
   4. \(\text{CCl}_4\)

7. The relatively high boiling point of water is due to water having

   1. hydrogen bonding
   2. metallic bonding
   3. nonpolar covalent bonding
   4. strong ionic bonding

8. Which statement explains why \(\text{H}_2\text{O}\) has a higher boiling point than \(\text{N}_2\)?

   1. \(\text{H}_2\text{O}\) has greater molar mass than \(\text{N}_2\).
   2. \(\text{H}_2\text{O}\) has less molar mass than \(\text{N}_2\).
   3. \(\text{H}_2\text{O}\) has stronger intermolecular forces than \(\text{N}_2\).
   4. \(\text{H}_2\text{O}\) has weaker intermolecular forces than \(\text{N}_2\).
9. The ability of carbon to attract electrons is
   1. greater than that of nitrogen, but less than that of oxygen
   2. less than that of nitrogen, but greater than that of oxygen
   3. greater than that of nitrogen and oxygen
   4. less than that of nitrogen and oxygen

10. Which statement explains why the radius of a lithium atom is larger than the radius of a lithium ion?
   1. Metals lose electrons when forming an ion.
   2. Metals gain electrons when forming an ion.
   3. Nonmetals lose electrons when forming an ion.
   4. Nonmetals gain electrons when forming an ion.

11. Compared to the radius of a chlorine atom, the radius of a chloride ion is
   1. larger because chlorine loses an electron
   2. larger because chlorine gains an electron
   3. smaller because chlorine loses an electron
   4. smaller because chlorine gains an electron

12. As an atom becomes an ion, its mass number
   1. decreases
   2. increases
   3. remains the same

13. Which of these formulas contains the most polar bond?
   1. H-Br
   2. A-Cl
   3. H-F
   4. H-I

14. Which element has atoms that can form single, double, and triple covalent bonds with other atoms of the same element?
   1. hydrogen
   2. oxygen
   3. fluorine
   4. carbon

15. Which change occurs when a barium atom loses two electrons?
   1. It becomes a negative ion and its radius decreases.
   2. It becomes a negative ion and its radius increases.
   3. It becomes a positive ion and its radius decreases.
   4. It becomes a positive ion and its radius increases.

16. As a chlorine atom becomes a negative ion, the atom
   1. gains an electron and its radius increases
   2. gains an electron and its radius decreases
   3. loses an electron and its radius increases
   4. loses an electron and its radius decreases
17. Based on Reference Table S, the atoms of which of these elements have the strongest attraction for electrons in a chemical bond?

1. N  
2. Na  
3. P  
4. Pt

18. Which symbol represents a particle that has the same total number of electrons as $S^{2-}$?

1. $O^{2-}$  
2. Si  
3. Se$^{2-}$  
4. Ar

19. Which element is malleable and can conduct electricity in the solid phase?

1. iodine  
2. phosphorus  
3. sulfur  
4. tin

20. Which type of bond results when one or more valence electrons are transferred from one atom to another?

1. a hydrogen bond  
2. an ionic bond  
3. a nonpolar covalent bond  
4. a polar covalent bond

21. What is the total number of electrons shared in the bonds between the two carbon atoms in a molecule of ... $\text{H–C}≡\text{C–H}$?

1. 6  
2. 2  
3. 3  
4. 8

22. Which formula represents a nonpolar molecule?

1. $\text{CH}_4$  
2. HCl  
3. $\text{H}_2\text{O}$  
4. $\text{NH}_3$

23. Which changes occur as a cadmium atom, Cd, becomes a cadmium ion, Cd$^{2+}$?

1. The Cd atom gains two electrons and its radius decreases.  
2. The Cd atom gains two electrons and its radius increases.  
3. The Cd atom loses two electrons and its radius decreases.  
4. The Cd atom loses two electrons and its radius increases.

24. Which element has atoms with the greatest attraction for electrons in a chemical bond?

1. beryllium  
2. fluorine  
3. lithium  
4. oxygen

25. What is the total number of electrons in a $S^{2-}$ ion?

1. 10  
2. 14  
3. 16  
4. 18
26. Given the key (see first diagram):
Which particle diagram represents a sample containing the compound CO(g)?

<table>
<thead>
<tr>
<th>Key</th>
</tr>
</thead>
<tbody>
<tr>
<td>○ = Atom of oxygen</td>
</tr>
<tr>
<td>● = Atom of carbon</td>
</tr>
</tbody>
</table>

(1)  (2)  (3)  (4)

27. When an atom loses one or more electrons, this atom becomes a
1. positive ion with a radius smaller than the radius of this atom
2. positive ion with a radius larger than the radius of this atom
3. negative ion with a radius smaller than the radius of this atom
4. negative ion with a radius larger than the radius of this atom

28. What is the name of the polyatomic ion in the compound Na₂O₂?
1. hydroxide
2. oxalate
3. oxide
4. peroxide

29. Given the balanced equation: I + I → I₂
Which statement describes the process represented by this equation?
1. A bond is formed as energy is absorbed.
2. A bond is formed and energy is released.
3. A bond is broken as energy is absorbed.
4. A bond is broken and energy is released.

30. An oxygen molecule contains a double bond because the two atoms of oxygen share a total of
1. 1 electron
2. 2 electrons
3. 3 electrons
4. 4 electrons

31. What is the total number of electrons in a Mg²⁺ ion?
1. 10
2. 12
3. 14
4. 24
32. Which element has an atom with the greatest attraction for electrons in a chemical bond?
   1. As                        3. N
   2. Bi                        4. P

33. As a bond between a hydrogen atom and a sulfur atom is formed, electrons are
   1. shared to form an ionic bond  3. transferred to form an ionic bond
   2. shared to form a covalent bond 4. transferred to form a covalent bond

34. Which formula represents a polar molecule?
   1. Br₂                       3. CH₄
   2. CO₂                       4. NH₃

35. What can be concluded if an ion of an element is smaller than an atom of the same element?
   1. The ion is negatively charged because it has fewer electrons than the atom.
   2. The ion is negatively charged because it has more electrons than the atom.
   3. The ion is positively charged because it has fewer electrons than the atom.
   4. The ion is positively charged because it has more electrons than the atom.

36. A barium atom attains a stable electron configuration when it bonds with
   1. one chlorine atom           3. one sodium atom
   2. two chlorine atoms          4. two sodium atoms

37. Which compound contains both ionic and covalent bonds?
   1. ammonia                    3. sodium nitrate
   2. methane                    4. potassium chloride

38. An atom in the ground state has a stable valence electron configuration. This atom could be an atom of
   1. Al                         3. Na
   2. Cl                         4. Ne

39. An atom of an element has a total of 12 electrons. An ion of the same element has a total of 10 electrons. Which statement describes the charge and radius of the ion?
   1. The ion is positively charged and its radius is smaller than the radius of the atom.
   2. The ion is positively charged and its radius is larger than the radius of the atom.
   3. The ion is negatively charged and its radius is smaller than the radius of the atom.
   4. The ion is negatively charged and its radius is larger than the radius of the atom.

40. Which formula represents a nonpolar molecule?
   1. CH₄                        3. H₂O
   2. HCl                        4. NH₃
41. The chemical bonding in sodium phosphate, Na$_3$PO$_4$, is classified as
   1. ionic, only
   2. metallic, only
   3. both covalent and ionic
   4. both covalent and metallic

42. Which element is composed of molecules that each contain a multiple covalent bond?
   1. chlorine
   2. fluorine
   3. hydrogen
   4. nitrogen

43. An atom of which element has the greatest attraction for electrons in a chemical bond?
   1. As
   2. Ga
   3. Ge
   4. Se

44. Which formula represents a polar molecule?
   1. H$_2$
   2. H$_2$O
   3. CO$_2$
   4. CCl$_4$

45. Two categories of compounds are
   1. covalent and molecular
   2. covalent and metallic
   3. ionic and molecular
   4. ionic and metallic

46. Which type of bond is found between atoms of solid cobalt?
   1. nonpolar covalent
   2. polar covalent
   3. metallic
   4. ionic

47. Which formula represents a molecule having a nonpolar covalent bond?

48. Given the balanced equation representing a reaction: Cl$_2$ → Cl + Cl
   What occurs during this reaction?
   1. A bond is broken as energy is absorbed.
   2. A bond is broken as energy is released.
   3. A bond is formed as energy is absorbed.
   4. A bond is formed as energy is released.
49. Which atom has the *weakest* attraction for the electrons in a bond with an H atom?
   1. Cl atom  
   2. F atom  
   3. O atom  
   4. S atom

50. Which statement explains why Br₂ is a liquid at STP and I₂ is a solid at STP?
   1. Molecules of Br₂ are polar, and molecules of I₂ are nonpolar.  
   2. Molecules of I₂ are polar, and molecules of Br₂ are nonpolar.  
   3. Molecules of Br₂ have stronger intermolecular forces than molecules of I₂.  
   4. Molecules of I₂ have stronger intermolecular forces than molecules of Br₂.

51. Which electron configuration is correct for a sodium ion?
   1. 2-7  
   2. 2-8  
   3. 2-8-1  
   4. 2-8-2

52. Which Lewis electron-dot structure is drawn correctly for the atom it represents?
   (1) :N  
   (2) :F  
   (3) :O  
   (4) :Ne

53. The bonds in the compound MgSO₄ can be described as
   1. ionic, only  
   2. covalent, only  
   3. both ionic and covalent  
   4. neither ionic nor covalent

54. Which particle has the same electron configuration as a potassium ion?
   1. fluoride ion  
   2. sodium ion  
   3. neon atom  
   4. argon atom

55. Which characteristic is a property of molecular substances?
   1. good heat conductivity  
   2. good electrical conductivity  
   3. low melting point  
   4. high melting point

56. Given the Lewis electron-dot diagram:
   Which electrons are represented by all of the dots?
   H
   H : C : H
   1. the carbon valence electrons, only  
   2. the hydrogen valence electrons, only  
   3. the carbon and hydrogen valence electrons  
   4. all of the carbon and hydrogen electrons
57. Based on intermolecular forces, which of these substances would have the highest boiling point?
   1. He  
   2. O₂  
   3. CH₄  
   4. NH₃

58. Which type of bonding is found in all molecular substances?
   1. covalent bonding  
   2. hydrogen bonding  
   3. ionic bonding  
   4. metallic bonding

59. What is the total number of electrons shared in a double covalent bond between two atoms?
   1. 1  
   2. 2  
   3. 8  
   4. 4

60. Which formula represents a nonpolar molecule?
   1. H₂S  
   2. HCl  
   3. CH₄  
   4. NH₃

61. What occurs when an atom loses an electron?
   1. The atom's radius decreases and the atom becomes a negative ion.  
   2. The atom's radius decreases and the atom becomes a positive ion.  
   3. The atom's radius increases and the atom becomes a negative ion.  
   4. The atom's radius increases and the atom becomes a positive ion.

62. Based on Reference Table S, atoms of which of these elements have the strongest attraction for the electrons in a chemical bond?
   1. Al  
   2. Si  
   3. P  
   4. S

63. Which Lewis electron-dot diagram in the image shown is correct for a S²⁻ ion?
   1. \[\text{[\:\cdot\:\cdot\:]}^2^-\]  
   2. \[\text{[\:\cdot\:]}^2^-\]  
   3. \[\text{[\:\cdot\:\cdot\:]}^2^-\]  
   4. \[\text{[\:\cdot\:\cdot\:]}^2^-\]

64. Which substance contains bonds that involved the transfer of electrons from one atom to another?
   1. CO₂  
   2. NH₃  
   3. KBr  
   4. Cl₂
65. What is the total number of pairs of electrons shared in a molecule of \( \text{N}_2 \)?
   1. one pair
   2. two pairs
   3. three pairs
   4. four pairs

66. Which formula represents a nonpolar molecule containing polar covalent bonds?
   1. \( \text{H}_2\text{O} \)
   2. \( \text{CCl}_4 \)
   3. \( \text{NH}_3 \)
   4. \( \text{H}_2 \)

67. The degree of polarity of a chemical bond in a molecule of a compound can be predicted by determining the difference in the
   1. melting points of the elements in the compound
   2. densities of the elements in the compound
   3. electronegativities of the bonded atoms in a molecule of the compound
   4. atomic masses of the bonded atoms in a molecule of the compound

68. Based on electronegativity values, which type of elements tends to have the greatest attraction for electrons in a bond?
   1. metals
   2. metalloids
   3. nonmetals
   4. noble gases

69. Atoms of which element have the greatest tendency to gain electrons?
   1. bromine
   2. chlorine
   3. fluorine
   4. iodine

70. Which polyatomic ion contains the greatest number of oxygen atoms?
   1. acetate
   2. carbonate
   3. hydroxide
   4. peroxide

71. Which formula represents an ionic compound?
   1. \( \text{H}_2 \)
   2. \( \text{CH}_4 \)
   3. \( \text{CH}_3\text{OH} \)
   4. \( \text{NH}_4\text{Cl} \)

72. An ion of which element has a larger radius than an atom of the same element?
   1. aluminum
   2. chlorine
   3. magnesium
   4. sodium

73. What is the total number of different elements present in \( \text{NH}_4\text{NO}_3 \)?
   1. 7
   2. 9
   3. 3
   4. 4
74. Given the balanced equation representing a reaction: \( \text{Cl}_2(g) \rightarrow \text{Cl}(g) + \text{Cl}(g) \)
What occurs during this change?

1. Energy is absorbed and a bond is broken. 3. Energy is released and a bond is broken.
2. Energy is absorbed and a bond is formed. 4. Energy is released and a bond is formed.

75. What is the net charge on an ion that has 9 protons, 11 neutrons, and 10 electrons?

1. 1+ 3. 1-
2. 2+ 4. 2-

76. At standard pressure, a certain compound has a low boiling point and is insoluble in water. At STP, this compound most likely exists as

1. ionic crystals 3. nonpolar molecules
2. metallic crystals 4. polar molecules

77. What is the total number of shared electrons in a molecule of this substance?

\[
\begin{align*}
\text{H} & \quad \text{H} & \quad \text{C} & \quad \text{C} & \quad \text{H} \\
\text{C} & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{H}
\end{align*}
\]

1. 22 3. 9
2. 11 4. 6

78. An atom of which element has the greatest attraction for the electrons in a bond with a hydrogen atom?

1. chlorine 3. silicon
2. phosphorus 4. sulfur

79. Which property could be used to identify a compound in the laboratory?

1. mass 3. temperature
2. melting point 4. volume

80. Which statement describes what occurs as two atoms of bromine combine to become a molecule of bromine?

1. Energy is absorbed as a bond is formed. 3. Energy is released as a bond is formed.
2. Energy is absorbed as a bond is broken. 4. Energy is released as a bond is broken.

81. What is the total number of electrons shared between the atoms represented in this formula?

\[
\begin{align*}
\text{O} & \quad \text{O} \\
\end{align*}
\]

1. 1 3. 8
2. 2 4. 4
82. Which formulas represent two polar molecules?
   1. CO₂ and HCl          3. H₂O and HCl
   2. CO₂ and CH₄          4. H₂O and CH₄

83. What is the total number of valence electrons in a sulfide ion in the ground state?
   1. 8          3. 16
   2. 2          4. 18

84. Which type of substance can conduct electricity in the liquid phase but not in the solid phase?
   1. ionic compound 3. metallic element
   2. molecular compound 4. nonmetallic element

85. Why is a molecule of CO₂ nonpolar even though the bonds between the carbon atom and the oxygen atoms are polar?
   1. The shape of the CO₂ molecule is symmetrical.
   2. The shape of the CO₂ molecule is asymmetrical.
   3. The CO₂ molecule has a deficiency of electrons.
   4. The CO₂ molecule has an excess of electrons.

86. Which formula represents a molecular compound?
   1. HI          3. KCl
   2. KI          4. LiCl

87. Which Group 15 element exists as diatomic molecules at STP?
   1. phosphorus 3. bismuth
   2. nitrogen 4. arsenic

88. What is the total number of electrons shared in a double covalent bond?
   1. 1          3. 3
   2. 2          4. 4

89. Given the balanced equation representing a reaction in the accompanying diagram:
   Br₂ + energy → Br + Br
Which statement describes the energy change and bonds in this reaction?
   1. Energy is released as bonds are broken.
   2. Energy is released as bonds are formed.
   3. Energy is absorbed as bonds are broken.
   4. Energy is absorbed as bonds are formed.

90. The bond between which two atoms is most polar?
   1. Br and Cl          3. I and Cl
   2. Br and F          4. I and F
91. In the formula \(X_2(SO_4)_3\), the \(X\) represents a metal. This metal could be located on the Periodic Table in

1. Group 1
2. Group 2
3. Group 13
4. Group 14

92. Which element forms an ionic compound when it reacts with lithium?

1. K
2. Fe
3. Kr
4. Br

93. Given the accompanying formula representing a molecule: \(H - C \equiv C - H\)

The molecule is

1. symmetrical and polar
2. symmetrical and nonpolar
3. asymmetrical and polar
4. asymmetrical and nonpolar

94. Which compound has both ionic and covalent bonds?

1. \(CO_2\)
2. \(CH_3OH\)
3. \(NaI\)
4. \(Na_2CO_3\)

95. An oxide ion (\(O^{2-}\)) formed from an oxygen-18 atom contains exactly

1. 8 protons, 8 neutrons, 10 electrons
2. 8 protons, 10 neutrons, 8 electrons
3. 8 protons, 10 neutrons, 10 electrons
4. 10 protons, 8 neutrons, 8 electrons

96. An unknown substance, liquid \(X\), is tested in the laboratory. The chemical and physical test results are listed below.

- Nonconductor of electricity
- Insoluble in water
- Soluble in hexane
- Low melting point as a solid
- Combustion produces only \(CO_2\) and \(H_2O\)

Based on these results, a student should conclude that liquid \(X\) is

1. ionic and organic
2. ionic and inorganic
3. covalent and organic
4. covalent and inorganic

97. In the box shown, or on a separate piece of paper, draw the electron-dot (Lewis) structure of calcium chloride. [2]
98. Base your answer on the information below.
Testing of an unknown solid shows that it has the properties listed below.
   (1) low melting point
   (2) nearly insoluble in water
   (3) nonconductor of electricity
   (4) relatively soft solid
State the type of bonding that would be expected in the particles of this substance. [1]

99. Testing of an unknown solid shows that it has the properties listed below.
   (1) low melting point
   (2) nearly insoluble in water
   (3) nonconductor of electricity
   (4) relatively soft solid
Explain why the particles of this substance are nonconductors of electricity. [1]

100. As a neutral sulfur atom gains two electrons, what happens to the radius of the atom? [1]

101. After a neutral sulfur atom gains two electrons, what is the resulting charge of the ion? [1]

102. Base your answer on the information below.
A safe level of fluoride ions is added to many public drinking water supplies. Fluoride ions have been found to help prevent tooth decay. Another common source of fluoride ions is toothpaste. One of the fluoride compounds used in toothpaste is tin(II) fluoride.
A town located downstream from a chemical plant was concerned about fluoride ions from the plant leaking into its drinking water. According to the Environmental Protection Agency, the fluoride ion concentration in drinking water cannot exceed 4 ppm. The town hired a chemist to analyze its water. The chemist determined that a 175-gram sample of the town's water contains 0.000 250 gram of fluoride ions.
In the box provided (print this page), draw a Lewis electron-dot diagram for a fluoride ion. [1]

103. Base your answer on the balanced equation: \(2\text{Na(s)} + \text{Cl}_2(g) \rightarrow 2\text{NaCl(s)}\)
    On a separate piece of paper, draw a Lewis electron-dot diagram for a molecule of chlorine, \(\text{Cl}_2\).

104. Base your answer on the balanced equation: \(2\text{Na(s)} + \text{Cl}_2(g) \rightarrow 2\text{NaCl(s)}\)
    Explain, in terms of electrons, why the bonding in \(\text{NaCl}\) is ionic.
Aluminum is one of the most abundant metals in Earth's crust. The aluminum compound found in bauxite ore is \( \text{Al}_2\text{O}_3 \). Over one hundred years ago, it was difficult and expensive to isolate aluminum from bauxite ore. In 1886, a brother and sister team, Charles and Julia Hall, found that molten (melted) cryolite, \( \text{Na}_3\text{AlF}_6 \), would dissolve bauxite ore. Electrolysis of the resulting mixture caused the aluminum ions in the \( \text{Al}_2\text{O}_3 \) to be reduced to molten aluminum metal. This less expensive process is known as the Hall process.

Explain, in terms of ions, why molten cryolite conducts electricity.

Explain, in terms of electronegativity, why a P–Cl bond in a molecule of \( \text{PCl}_5 \) is more polar than a P–S bond in a molecule of \( \text{P}_2\text{S}_5 \).

A 1.00-mole sample of neon gas occupies a volume of 24.4 liters at 298 K and 101.3 kilopascals. In the space provided or on a separate piece of paper, calculate the density of this sample. Your response must include both a correct numerical setup and the calculated results.

The hydrocarbon 2-methylpropane reacts with iodine as represented by the balanced equation below. At standard pressure, the boiling point of 2-methylpropane is lower than the boiling point of 2-iodo-2-methylpropane. Explain the difference in the boiling points of 2-methylpropane and 2-iodo-2-methylpropane in terms of both molecular polarity and intermolecular forces.

An unlit candle is secured to the bottom of a 200-milliliter glass beaker. Baking soda (sodium hydrogen carbonate) is added around the base of the candle as shown below. The candle is lit and dilute ethanoic acid is poured down the inside of the beaker. As the acid reacts with the baking soda, bubbles of \( \text{CO}_2 \) gas form. After a few seconds, the air in the beaker is replaced by 0.20 liter of \( \text{CO}_2 \) gas, causing the candle flame to go out. The density of \( \text{CO}_2 \) gas is 1.8 grams per liter at room temperature. Calculate the mass of the \( \text{CO}_2 \) gas that replaced the air in the beaker. Your response must include both a correct numerical setup and the calculated result.

What is the total number of electron pairs shared between the carbon atom and one of the oxygen atoms in a carbon dioxide molecule?

Explain, in terms of subatomic particles, why the radius of a chloride ion is larger than the radius of a chlorine atom.

Explain, in terms of valence electrons, why the bonding in magnesium oxide, \( \text{MgO} \), is similar to the bonding in barium chloride, \( \text{BaCl}_2 \).

Base your answer on the information below. Natural gas is a mixture that includes butane, ethane, methane, and propane. Differences in boiling points can be used to separate the components of natural gas. The boiling points at standard pressure for these components are listed in the accompanying table. List the four components of natural gas in order of increasing strength of intermolecular forces.
114. On a separate sheet of paper, draw an electron-dot diagram for each of the following substances:
   a) calcium oxide (an ionic compound) [1]
   b) hydrogen bromide [1]
   c) carbon dioxide [1]

115. Base your answer on the article below, the Reference Tables for Physical Setting/Chemistry, and your knowledge of chemistry.
   In the 1920s, paint used to inscribe the numbers on watch dials was composed of a luminescent (glow-in-the-dark) mixture. The powdered-paint base was a mixture of radium salts and zinc sulfide. As the paint was mixed, the powdered base became airborne and drifted throughout the workroom causing the contents of the workroom, including the painters' clothes and bodies, to glow in the dark. The paint is luminescent because radiation from the radium salts strikes a scintillator. A scintillator is a material that emits visible light in response to ionizing radiation. In watchdial paint, zinc sulfide acts as the scintillator.
   Radium present in the radium salts decomposes spontaneously, emitting alpha particles. These particles can cause damage to the body when they enter human tissue. Alpha particles are especially harmful to the blood, liver, lungs, and spleen because they can alter genetic information in the cells. Radium can be deposited in the bones because it substitutes for calcium. On a separate sheet of paper explain why zinc sulfide is used in luminescent paint. [1]

116. Base your answer on the electronegativity values and atomic numbers of fluorine, chlorine, bromine, and iodine that are listed on Reference Table S.
   Explain, in terms of electronegativity, why the H--F bond is expected to be more polar than the H--I bond. [1]

117. Base your answer on the information shown.
   Potassium ions are essential to human health. The movement of dissolved potassium ions, K⁺(aq), in and out of a nerve cell allows that cell to transmit an electrical impulse.
   a) What is the total number of electrons in a potassium ion? [1]
   b) Explain, in terms of atomic structure, why a potassium ion is smaller than a potassium atom. [1]

118. Base your answer on the information shown.
   Potassium ions are essential to human health. The movement of dissolved potassium ions, K⁺(aq), in and out of a nerve cell allows that cell to transmit an electrical impulse.
   What property of potassium ions allows them to transmit an electrical impulse? [1]
119. In the space provided or on a separate piece of paper, draw a Lewis electron-dot diagram for a molecule of phosphorus trichloride, PCl₃.

120. Base your answer on the information below.
A metal, M, was obtained from a compound in a rock sample. Experiments have determined that the element is a member of Group 2 on the Periodic Table of the Elements.
a) Explain, in terms of electrons, why element M is a good conductor of electricity. [1]
b) Explain why the radius of a positive ion of element M is smaller than the radius of an atom of element M. [1]
c) Using the symbol M for the element, write the chemical formula for the compound that forms when element M reacts with iodine. [1]

121. Base your answer on the information below.
Elements with atomic numbers 112 and 114 have been produced and their IUPAC names are pending approval. However, an element that would be put between these two elements on the Periodic Table has not yet been produced. If produced, this element will be identified by the symbol Uut until an IUPAC name is approved.
In the space on the answer sheet, or, if taken online, on a separate piece of paper, draw a Lewis electron-dot diagram for an atom of Uut. [1]

122. Base your answer on the information below.
Rust on an automobile door contains Fe₂O₃(s). The balanced equation representing one of the reactions between iron in the door of the automobile and oxygen in the atmosphere is given below.
4Fe(s) + 3O₂(g) → 2Fe₂O₃(s)
Write the IUPAC name for Fe₂O₃. [1]

123. Describe one appropriate laboratory test that can be used to determine the malleability of a solid sample of an element at room temperature. [1]

124. Base your answer on the information below.
A piece of magnesium ribbon is reacted with excess hydrochloric acid to produce aqueous magnesium chloride and hydrogen gas. The volume of the dry hydrogen gas produced is 45.6 milliliters. The temperature of the gas is 293 K, and the pressure is 99.5 kilopascals.
Identify the type of bond between the atoms in a molecule of the gas produced in this laboratory investigation. [1]

125. Base your answer on the information below.
A portable propane-fueled lantern contains a mesh silk bag coated with metal hydroxides. The primary metal hydroxide is yttrium hydroxide. When the silk bag is installed, it is ignited and burned away, leaving the metal hydroxide coating. The coating forms metal oxides that glow brightly when heated to a high temperature. During a test, a propane lantern is operated for three hours and consumes 5.0 moles of propane from the lantern's tank. The balanced equation below represents the combustion of propane.
C₃H₈ + 5O₂ → 3CO₂ + 4H₂O + energy
Write the formula for the primary metal hydroxide used in the lantern. [1]
126. Base your answer on the information below.
When a person perspires (sweats), the body loses many sodium ions and potassium ions. The evaporation of sweat cools the skin.
After a strenuous workout, people often quench their thirst with sports drinks that contain NaCl and KCl. A single 250-gram serving of one sports drink contains 0.055 gram of sodium ions. In the space on the answer sheet or on a separate piece of paper, draw a Lewis electron-dot diagram for one of the positive ions lost by the body as a person perspires. [1]

127. Base your answer on the information below and the accompanying table.
Bond energy is the amount of energy required to break a chemical bond. The table gives a formula and the carbon-nitrogen bond energy for selected nitrogen compounds.
a) Describe, in terms of electrons, the type of bonding between the carbon atom and the nitrogen atom in a molecule of methanamine. [1]
b) Identify the noble gas that has atoms in the ground state with the same electron configuration as the nitrogen in a molecule of isocyanic acid. [1]
c) Explain, in terms of charge distribution, why a molecule of hydrogen cyanide is polar. [1]

<table>
<thead>
<tr>
<th>Selected Nitrogen Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Compound</strong></td>
</tr>
<tr>
<td>hydrogen cyanide</td>
</tr>
<tr>
<td>isocyanic acid</td>
</tr>
</tbody>
</table>
| methanamine | H

Laws of conservation:

1. In a redox reaction, there is conservation of
   1. mass, only  
   2. charge, only  
   3. both mass and charge  
   4. neither mass nor charge

2. In this reaction there is conservation of
   \( \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(l) + \text{energy} \)
   1. mass, only  
   2. mass and charge, only  
   3. charge and energy, only  
   4. charge, energy, and mass

3. Given the balanced equation representing a reaction: \( 2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \)
   What is the total mass of water formed when 8 grams of hydrogen reacts completely with 64 grams of oxygen?
   1. 18 g  
   2. 36 g  
   3. 56 g  
   4. 72 g

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4. In a chemical reaction, there is conservation of
   1. energy, volume, and mass
   2. energy, volume, and charge
   3. mass, charge, and energy
   4. mass, charge, and volume

5. Given the balanced equation representing a reaction: \(2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}\)
   What is the mass of \(\text{H}_2\text{O}\) produced when 10.0 grams of \(\text{H}_2\) reacts completely with 80.0 grams of \(\text{O}_2\)?
   1. 70.0 g
   2. 90.0 g
   3. 180. g
   4. 800. g

6. Given the balanced ionic equation: \(2\text{Al}(s) + 3\text{Cu}^{2+}(\text{aq}) \rightarrow 2\text{Al}^{3+}(\text{aq}) + 3\text{Cu}(s)\)
   Compared to the total charge of the reactants, the total charge of the products is
   1. less
   2. greater
   3. the same

7. Given the balanced equation representing a reaction:
   \(\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(l) + 55.8 \text{ kJ}\)
   In this reaction there is conservation of
   1. mass, only
   2. mass and charge, only
   3. mass and energy, only
   4. mass, charge, and energy

8. Which equation shows conservation of mass and charge?
   1. \(\text{NH}_4\text{Br} \rightarrow \text{NH}_3 + \text{Br}_2\)
   2. \(2\text{Mg} + \text{Fe}^{3+} \rightarrow \text{Mg}^{2+} + 3\text{Fe}\)
   3. \(\text{H}_2\text{SO}_4 + \text{LiOH} \rightarrow \text{Li}_2\text{SO}_4 + \text{H}_2\text{O}\)
   4. \(\text{Cu} + 2\text{Ag}^+ \rightarrow \text{Cu}^{2+} + 2\text{Ag}\)

9. A balanced equation representing a chemical reaction can be written using
   1. chemical formulas and mass numbers
   2. chemical formulas and coefficients
   3. first ionization energies and mass numbers
   4. first ionization energies and coefficients

10. Given the balanced equation representing a reaction: \(2\text{NaCl}(l) \rightarrow 2\text{Na}(l) + \text{Cl}_2(g)\)
    A 1170.-gram sample of \(\text{NaCl}(l)\) completely reacts, producing 460. grams of \(\text{Na}(l)\). What is the total mass of \(\text{Cl}_2(g)\) produced?
    1. 355 g
    2. 710. g
    3. 1420. g
    4. 1630. g

11. Which equation shows conservation of charge?
    (1) \(\text{Fe} \rightarrow \text{Fe}^{2+} + \text{e}^-\)
    (2) \(\text{Fe} + 2\text{e}^- \rightarrow \text{Fe}^{2+}\)
    (3) \(\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-\)
    (4) \(\text{Fe} + 2\text{e}^- \rightarrow \text{Fe}^{3+}\)
12. Given the reaction between two different elements in the gaseous state: (see image) Box A represents a mixture of the two reactants before the reaction occurs. The product of this reaction is a gas. On a separated piece of paper, draw the system after the reaction has gone to completion, based on the Law of Conservation of Matter. [2]

\[ \text{○○○} + \text{●●●} \rightarrow \text{●●�} + \text{●●○} \]

13. Base your answer on the balanced chemical equation: \( 2 \text{H}_2\text{O} \rightarrow 2 \text{H}_2 + \text{O}_2 \)
How does the balanced chemical equation show the Law of Conservation of Mass? [1]

14. Arsenic is often obtained by heating the ore arsenopyrite, FeAsS. The decomposition of FeAsS is represented by the balanced equation shown.
In the solid phase, arsenic occurs in two forms. One form, yellow arsenic, has a density of 1.97 g/cm\(^3\) at STP. The other form, gray arsenic, has a density of 5.78 g/cm\(^3\) at STP. When arsenic is heated rapidly in air, arsenic(III) oxide is formed.
Although arsenic is toxic, it is needed by the human body in very small amounts. The body of a healthy human adult contains approximately 5 milligrams of arsenic.
When heated, a 125.0-kilogram sample of arsenopyrite yields 67.5 kilograms of FeS. Determine the total mass of arsenic produced in this reaction. [1]

\[ \text{FeAsS}(s) \xrightarrow{\text{heat}} \text{FeS}(s) + \text{As(g)} \]

15. The Solvay process is a multistep industrial process used to produce washing soda, Na\(_2\)CO\(_3\)(s). In the last step of the Solvay process, NaHCO\(_3\)(s) is heated to 300°C, producing washing soda, water, and carbon dioxide. This reaction is represented by the balanced equation below.

\[ 2\text{NaHCO}_3(s) + \text{heat} \rightarrow \text{Na}_2\text{CO}_3(s) + \text{H}_2\text{O}(g) + \text{CO}_2(g) \]
Determine the total mass of washing soda produced if 3360 kilograms of NaHCO\(_3\) reacts completely to produce 360 kilograms of H\(_2\)O and 880 kilograms of CO\(_2\). [1]

16. Base your answer on the information below.
Hydrogen peroxide, H\(_2\)O\(_2\), is a water-soluble compound. The concentration of an aqueous hydrogen peroxide solution that is 3% by mass H\(_2\)O\(_2\) is used as an antiseptic. When the solution is poured on a small cut in the skin, H\(_2\)O\(_2\) reacts according to the balanced equation below.

\[ 2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2 \]
Calculate the total mass of H\(_2\)O\(_2\) in 20.0 grams of an aqueous H\(_2\)O\(_2\) solution that is used as an antiseptic. Your response must include both a numerical setup and the calculated result. [2]
**Percent by mass:**

1. A sample of a substance containing only magnesium and chlorine was tested in the laboratory and was found to be composed of 74.5% chlorine by mass. If the total mass of the sample was 190.2 grams, what was the mass of the magnesium?
   1. 24.3 g
   2. 48.5 g
   3. 70.9 g
   4. 142 g

2. A hydrated salt is a solid that includes water molecules within its crystal structure. A student heated a 9.10-gram sample of a hydrated salt to a constant mass of 5.41 grams. What percent by mass of water did the salt contain?
   1. 3.69%
   2. 16.8%
   3. 40.5%
   4. 59.5%

3. In which compound is the percent composition by mass of chlorine equal to 42%?
   1. HClO (gram-formula mass = 52 g/mol)
   2. HClO₂ (gram-formula mass = 68 g/mol)
   3. HClO₃ (gram-formula mass = 84 g/mol)
   4. HClO₄ (gram-formula mass = 100. g/mol)

4. What is the percent composition by mass of sulfur in the compound MgSO₄ (gram-formula mass = 120. grams per mole)?
   1. 20.%
   2. 27%
   3. 46%
   4. 53%

5. Note: This question may require the use of the Reference Tables for Physical Setting/Chemistry. The percentage by mass of Br in the compound AlBr₃ is closest to
   1. 10.%
   2. 25%
   3. 75%
   4. 90.%

6. The percent composition by mass of magnesium in MgBr₂ (gram-formula mass = 184 grams/mole) is equal to
   1. 24/184 x 100
   2. 160/184 x 100
   3. 184/24 x 100
   4. 184/160 x 100

7. What is the percent composition by mass of hydrogen in NH₄HCO₃ (gram-formula mass = 79 grams/mole)?
   1. 5.1%
   2. 6.3%
   3. 10.%
   4. 50.%

8. The percent composition by mass of nitrogen in NH₄OH (gram-formula mass = 35 grams/mole) is equal to
   
   \[
   \begin{align*}
   (1) \quad & \frac{4}{35} \times 100 \\
   (2) \quad & \frac{14}{35} \times 100 \\
   (3) \quad & \frac{35}{14} \times 100 \\
   (4) \quad & \frac{35}{4} \times 100
   \end{align*}
   \]
9. Determine the percent composition by mass of oxygen in the compound $C_6H_{12}O_6$. [1]

10. Base your answer on the information below.
Arsenic is often obtained by heating the ore arsenopyrite, FeAsS. The decomposition of FeAsS is represented by the balanced equation shown.

In the solid phase, arsenic occurs in two forms. One form, yellow arsenic, has a density of 1.97 g/cm$^3$ at STP. The other form, gray arsenic, has a density of 5.78 g/cm$^3$ at STP. When arsenic is heated rapidly in air, arsenic(III) oxide is formed.
Although arsenic is toxic, it is needed by the human body in very small amounts. The body of a healthy human adult contains approximately 5 milligrams of arsenic.
Calculate the percent composition by mass of arsenic in arsenopyrite. Your response must include both a correct numerical setup and the calculated result. [2]

$$\text{FeAsS(s)} \xrightarrow{\text{heat}} \text{FeS(s)} + \text{As(g)}$$

11. Base your answer on the information below.
When a person perspires (sweats), the body loses many sodium ions and potassium ions. The evaporation of sweat cools the skin.
After a strenuous workout, people often quench their thirst with sports drinks that contain NaCl and KCl. A single 250-gram serving of one sports drink contains 0.055 gram of sodium ions.
In the space on the answer sheet or on a separate piece of paper, show a correct numerical setup for calculating the concentration of sodium ions in this sports drink, expressed as percent by mass. [1]

Compounds, naming compounds and writing formulas:

1. The chemical formula for nickel (II) bromide is
   1. Ni$_2$Br
   2. NiBr$_2$
   3. N$_2$Br
   4. NBr$_2$

2. What is the formula of titanium(II) oxide?
   1. TiO
   2. TiO$_2$
   3. Ti$_2$O
   4. Ti$_2$O$_3$

3. Which substance can be decomposed by a chemical change?
   1. calcium
   2. potassium
   3. copper
   4. ammonia

4. What is the IUPAC name for the compound FeS?
   1. iron(II) sulfate
   2. iron(III) sulfate
   3. iron(II) sulfide
   4. iron(III) sulfide

5. Which substance can be decomposed by chemical means?
   1. ammonia
   2. oxygen
   3. phosphorus
   4. silicon
6. Which formula represents lead(II) chromate?
1. PbCrO₄ 3. Pb₂CrO₄
2. Pb(CrO₄)₂ 4. Pb₂(CrO₄)₃

7. Which statement describes the composition of potassium chlorate, KClO₃?
1. The proportion by mass of elements combined in potassium chlorate is fixed.
3. Potassium chlorate is composed of four elements.
2. The proportion by mass of elements combined in potassium chlorate varies.
4. Potassium chlorate is composed of five elements.

8. Which substance can be decomposed by chemical means?
1. aluminum 3. silicon
2. octane 4. xenon

9. What is the chemical formula of iron(III) sulfide?
1. FeS 3. FeSO₃
2. Fe₂S₃ 4. Fe₂(SO₃)₃

10. What is the correct IUPAC name for the compound NH₄Cl?
1. nitrogen chloride 3. ammonium chloride
2. nitrogen chlorate 4. ammonium chlorate

11. Two different samples decompose when heated. Only one of the samples is soluble in water. Based on this information, these two samples are
1. both the same element 3. both the same compound
2. two different elements 4. two different compounds

12. The correct chemical formula for iron(II) sulfide is
1. FeS 3. FeSO₄
2. Fe₂S₃ 4. Fe₂(SO₄)₃

13. A compound is made up of iron and oxygen, only. The ratio of iron ions to oxide ions is 2:3 in this compound. The IUPAC name for this compound is
1. triiron dioxide 3. iron(III) oxide
2. iron (II) oxide 4. iron trioxide

14. Matter that is composed of two or more different elements chemically combined in a fixed proportion is classified as
1. a compound 3. a mixture
2. an isotope 4. a solution
15. Which list of formulas represents compounds, only?
   1. CO₂, H₂O, NH₃
   2. H₂, N₂, O₂
   3. H², Ne, NaCl
   4. MgO, NaCl, O₂

16. What is the chemical formula for iron(III) oxide?
   1. FeO
   2. Fe₂O₃
   3. Fe₃O
   4. Fe₃O₂

17. Which type of matter is composed of two or more elements that are chemically combined in a fixed proportion?
   1. solution
   2. compound
   3. homogeneous mixture
   4. heterogeneous mixture

18. Which formula represents strontium phosphate?
   1. SrPO₄
   2. Sr₃PO₈
   3. Sr₂(PO₄)₃
   4. Sr₃(PO₄)₂

19. Every water molecule has two hydrogen atoms bonded to one oxygen atom. This fact supports the concept that elements in a compound are
   1. chemically combined in a fixed proportion
   2. chemically combined in proportions that vary
   3. physically mixed in a fixed proportion
   4. physically mixed in proportions that vary

20. What is the formula of nitrogen (II) oxide?
   1. NO
   2. NO₂
   3. N₂O
   4. N₂O₄

21. Base your answer on the information and equation below.
    Antacids can be used to neutralize excess stomach acid. Brand A antacid contains the acid-neutralizing agent magnesium hydroxide, Mg(OH)₂. It reacts with HCl(aq) in the stomach, according to the following balanced equation: 2 HCl(aq) + Mg(OH)₂(s) → MgCl₂(aq) + 2 H₂O(l)
    Brand B antacid contains the acid-neutralizing agent sodium hydrogen carbonate. Write the chemical formula for sodium hydrogen carbonate. [1]

22. Base your answer on the information below.
    A safe level of fluoride ions is added to many public drinking water supplies. Fluoride ions have been found to help prevent tooth decay. Another common source of fluoride ions is toothpaste. One of the fluoride compounds used in toothpaste is tin(II) fluoride.
    A town located downstream from a chemical plant was concerned about fluoride ions from the plant leaking into its drinking water. According to the Environmental Protection Agency, the fluoride ion concentration in drinking water cannot exceed 4 ppm. The town hired a chemist to analyze its water. The chemist determined that a 175-gram sample of the town's water contains 0.000 250 gram of fluoride ions.
    What is the chemical formula for tin(II) fluoride? [1]
23. Base your answer on the information below.
An unlit candle is secured to the bottom of a 200-milliliter glass beaker. Baking soda (sodium hydrogen carbonate) is added around the base of the candle as shown below. The candle is lit and dilute ethanoic acid is poured down the inside of the beaker. As the acid reacts with the baking soda, bubbles of CO\(_2\) gas form. After a few seconds, the air in the beaker is replaced by 0.20 liter of CO\(_2\) gas, causing the candle flame to go out. The density of CO\(_2\) gas is 1.8 grams per liter at room temperature.
Write the chemical formula for baking soda. [1]

24. Base your answer on the information below.
Arsenic is often obtained by heating the ore arsenopyrite, FeAsS. The decomposition of FeAsS is represented by the balanced equation shown.
In the solid phase, arsenic occurs in two forms. One form, yellow arsenic, has a density of 1.97 g/cm\(^3\) at STP. The other form, gray arsenic, has a density of 5.78 g/cm\(^3\) at STP. When arsenic is heated rapidly in air, arsenic(III) oxide is formed.
Although arsenic is toxic, it is needed by the human body in very small amounts. The body of a healthy human adult contains approximately 5 milligrams of arsenic.
Write the formula for the compound produced when arsenic is heated rapidly in air. [1]

\[
\text{FeAsS(s)} \xrightarrow{\text{heat}} \text{FeS(s)} + \text{As(g)}
\]

25. Base your answer on the information below.
The Solvay process is a multistep industrial process used to produce washing soda, Na\(_2\)CO\(_3\)(s). In the last step of the Solvay process, NaHCO\(_3\)(s) is heated to 300°C, producing washing soda, water, and carbon dioxide. This reaction is represented by the balanced equation below.
\[
2\text{NaHCO}_3(s) + \text{heat} \rightarrow \text{Na}_2\text{CO}_3(s) + \text{H}_2\text{O}(g) + \text{CO}_2(g)
\]
Write the IUPAC name for washing soda. [1]

26. Base your answer on the information below.
Two sources of copper are cuprite, which has the IUPAC name copper(I) oxide, and malachite, which has the formula Cu\(_2\)CO\(_3\)(OH)\(_2\). Copper is used in home wiring and electric motors because it has good electrical conductivity. Other uses of copper not related to its electrical conductivity include coins, plumbing, roofing, and cooking pans. Aluminum is also used for cooking pans.
At room temperature, the electrical conductivity of a copper wire is 1.6 times greater than an aluminum wire with the same length and cross-sectional area. At room temperature, the heat conductivity of copper is 1.8 times greater than the heat conductivity of aluminum. At STP, the density of copper is 3.3 times greater than the density of aluminum.
Write the chemical formula of cuprite. [1]
Chemical reactions and types of reactions:

1. Given the lead-acid battery reaction: $\text{Pb}(s) + \text{PbO}_2(s) + 2\text{H}_2\text{SO}_4(\text{aq}) \underset{\text{charge}}{\overset{\text{discharge}}{\rightleftharpoons}} 2\text{PbSO}_4(s) + 2\text{H}_2\text{O}(\ell)$

   1. reactant, with decreasing concentration 3. product, with decreasing concentration
   2. reactant, with increasing concentration 4. product, with increasing concentration

2. Which list includes three types of chemical reactions?
   1. condensation, double replacement, and sublimation
   2. condensation, solidification, and synthesis
   3. decomposition, double replacement, and synthesis
   4. decomposition, solidification, and sublimation

3. Given the balanced equation representing a reaction: $4\text{Al}(s) + 3\text{O}_2(\text{g}) \rightarrow 2\text{Al}_2\text{O}_3(s)$

   Which type of chemical reaction is represented by this equation?
   1. double replacement
   2. single replacement
   3. substitution
   4. synthesis

4. Given the balanced equation representing a reaction: $\text{Zn}(s) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{H}_2(\text{g})$

   Which type of reaction is represented by this equation?
   1. decomposition
   2. double replacement
   3. single replacement
   4. synthesis

5. Note: This question may require the use of the Reference Tables for Physical Setting/Chemistry.
   Given the balanced equation: $\text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{NaNO}_3(\text{aq}) + \text{AgCl}(s)$

   This reaction is classified as
   1. synthesis
   2. decomposition
   3. single replacement
   4. double replacement

6. Which equation represents a decomposition reaction?
   (1) $\text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(\text{g})$
   (2) $\text{Cu}(s) + 2\text{AgNO}_3(\text{aq}) \rightarrow 2\text{Ag}(s) + \text{Cu(NO}_3)_2(\text{aq})$
   (3) $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\ell)$
   (4) $\text{KOH}(\text{aq}) + \text{HCl}(\text{aq}) \rightarrow \text{KCl}(\text{aq}) + \text{H}_2\text{O}(\ell)$

7. Which two solutions, when mixed together, will undergo a double replacement reaction and form a white, solid substance?
   1. NaCl(aq) and LiNO$_3$(aq)
   2. KCl(aq) and AgNO$_3$(aq)
   3. KCl(aq) and LiCl(aq)
   4. NaNO$_3$(aq) and AgNO$_3$(aq)
8. Base your answer on the balanced chemical equation: \( 2 \text{H}_2\text{O} \rightarrow 2 \text{H}_2 + \text{O}_2 \)
What type of reaction does this equation represent? [1]

9. Base your answer on the information below.
The Solvay process is a multistep industrial process used to produce washing soda, \( \text{Na}_2\text{CO}_3 \text{(s)} \). In the last step of the Solvay process, \( \text{NaHCO}_3 \text{(s)} \) is heated to 300°C, producing washing soda, water, and carbon dioxide. This reaction is represented by the balanced equation below.
\[
2\text{NaHCO}_3\text{(s)} + \text{heat} \rightarrow \text{Na}_2\text{CO}_3\text{(s)} + \text{H}_2\text{O(g)} + \text{CO}_2\text{(g)}
\]
Identify the type of chemical reaction represented by the equation. [1]

10. Base your answer on the information below.
A student places a 2.50-gram sample of magnesium metal in a bottle and fits the bottle with a 2-hole stopper as shown in the accompanying diagram. Hydrochloric acid is added to the bottle, causing a reaction. As the reaction proceeds, hydrogen gas travels through the tubing to an inverted bottle filled with water, displacing some of the water in the bottle.
Identify the type of chemical reaction that occurs when magnesium reacts with hydrochloric acid. [1]

![Diagram of magnesium reacting with hydrochloric acid](image)

11. Base your answer on the information below.
Hydrogen peroxide, \( \text{H}_2\text{O}_2 \), is a water-soluble compound. The concentration of an aqueous hydrogen peroxide solution that is 3% by mass \( \text{H}_2\text{O}_2 \) is used as an antiseptic. When the solution is poured on a small cut in the skin, \( \text{H}_2\text{O}_2 \) reacts according to the balanced equation below.
\[
2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2
\]
Identify the type of chemical reaction represented by the balanced equation. [1]

**Balancing equations:**

1. Given the incomplete equation: \( 4\text{Fe} + 3\text{O}_2 \rightarrow 2X \)
Which compound is represented by \( X \)?
   1. \( \text{FeO} \)
   2. \( \text{Fe}_2\text{O}_3 \)
   3. \( \text{Fe}_3\text{O}_2 \)
   4. \( \text{Fe}_3\text{O}_4 \)
2. Acid rain is a problem in industrialized countries around the world. Oxides of sulfur and nitrogen are formed when various fuels are burned. These oxides dissolve in atmospheric water droplets that fall to earth as acid rain or acid snow. While normal rain has a pH between 5.0 and 6.0 due to the presence of dissolved carbon dioxide, acid rain often has a pH of 4.0 or lower. This level of acidity can damage trees and plants, leach minerals from the soil, and cause the death of aquatic animals and plants. If the pH of the soil is too low, then quicklime, CaO, can be added to the soil to increase the pH. Quicklime produces calcium hydroxide when it dissolves in water. Balance the accompanying neutralization equation on a separate piece of paper, using the smallest whole-number coefficients.

**Answer Sheet**

\[ \text{HNO}_3 + \text{Ca(OH)}_2 \rightarrow \text{Ca(NO}_3)_2 + \text{H}_2\text{O} \]

3. Air bags are an important safety feature in modern automobiles. An air bag is inflated in milliseconds by the explosive decomposition of NaN$_3$(s). The decomposition reaction produces N$_2$(g), as well as Na(s), according to the unbalanced equation below.

\[ \text{NaN}_3(s) \rightarrow \text{Na(s)} + \text{N}_2(g) \]

Balance the incomplete equation shown (on a separate piece of paper, if necessary) for the decomposition of NaN$_3$, using the smallest wholenumber coefficients.

\[ \text{NaN}_3(s) \rightarrow \text{Na(s)} + \text{N}_2(g) \]

4. The unbalanced equation below represents the decomposition of potassium chlorate. KClO$_3$(s) → KCl(s) + O$_2$(g)

\( a \) Balance the equation in the space provided or on a separate piece of paper, using the smallest whole-number coefficients. [1]

\( b \) Determine the oxidation number of chlorine in the reactant. [1]

\[ \text{KClO}_3(s) \rightarrow \text{KCl(s)} + \text{O}_2(g) \]

5. A flashlight can be powered by a rechargeable nickel-cadmium battery. In the battery, the anode is Cd(s) and the cathode is NiO$_2$(s). The unbalanced equation below represents the reaction that occurs as the battery produces electricity. When a nickel-cadmium battery is recharged, the reverse reaction occurs (see accompanying equation). Balance the equation for the reaction that produces electricity, using the smallest whole-number coefficients. [1]

\[ \text{Cd(s)} + \text{NiO}_2(s) + \text{H}_2\text{O}(\ell) \rightarrow \text{Cd(OH)}_2(s) + \text{Ni(OH)}_2(s) \]

6. Ozone, O$_3$(g), is produced from oxygen, O$_2$(g), by electrical discharge during thunderstorms. The unbalanced accompanying equation (see image) represents the reaction that forms ozone. Balance the equation on the answer sheet or on a separate piece of paper for the production of ozone, using the smallest whole-number coefficients. [1]

\[ \text{O}_3(g) \rightarrow \text{O}_2(g) \]
7. Base your answer on the information shown, which describes the smelting of iron ore, and on your knowledge of chemistry.
In the smelting of iron ore, Fe₂O₃ is reduced in a blast furnace at high temperature by a reaction with carbon monoxide. Crushed limestone, CaCO₃, is also added to the mixture to remove impurities in the ore. The carbon monoxide is formed by the oxidation of carbon (coke), as shown in the reaction shown:

\[ 2 \text{C} + \text{O}_2 \rightarrow 2 \text{CO} + \text{energy} \]

Liquid iron flows from the bottom of the blast furnace and is processed into different alloys of iron. Balance the equation for the reaction of Fe₂O₃ and CO on a separate piece of paper, using the smallest whole-number coefficients. [1]

8. The decomposition of sodium azide, NaN₃(s), is used to inflate airbags. On impact, the NaN₃(s) is ignited by an electrical spark, producing N₂(g) and Na(s). The N₂(g) inflates the airbag. Balance the equation shown (on a separate piece of paper if necessary), using the smallest whole-number coefficients.

\[ ____ \text{NaN}_3(s) \rightarrow ____ \text{Na}(s) \rightarrow ____ \text{N}_2(g) \]

9. A student places a 2.50-gram sample of magnesium metal in a bottle and fits the bottle with a 2-hole stopper as shown in the accompanying diagram. Hydrochloric acid is added to the bottle, causing a reaction. As the reaction proceeds, hydrogen gas travels through the tubing to an inverted bottle filled with water, displacing some of the water in the bottle. Balance the equation shown for the reaction of magnesium and hydrochloric acid, using the smallest whole-number coefficients. [1]

\[ ____ \text{Mg(s)} + ____ \text{HCl(aq)} \rightarrow ____ \text{MgCl}_2(aq) + ____ \text{H}_2(g) \]

10. Base your answer on the information below.
A 1.0-gram strip of zinc is reacted with hydrochloric acid in a test tube. The unbalanced equation below represents the reaction.

\[ \text{Zn(s)} + \text{HCl(aq)} \rightarrow \text{H}_2(g) + \text{ZnCl}_2(aq) \]

Balance the equation on the answer sheet for the reaction of zinc and hydrochloric acid, using the smallest whole-number coefficients. [1]

\[ ____ \text{Zn(s)} + ____ \text{HCl(aq)} \rightarrow ____ \text{H}_2(g) + ____ \text{ZnCl}_2(aq) \]

11. Skiers, snowmobilers, and others involved in outdoor winter recreation use disposable heat packs. These heat packs are porous paper pouches containing sawdust, powdered carbon, sodium chloride, powdered iron, and Zeolite®. During production, this mixture is moistened slightly with water and then sealed in an airtight plastic pack. The reaction starts when the pack is opened and the mixture is exposed to air.

Given the unbalanced equation: . . . (see image)
a) Balance the equation, using smallest whole number coefficients. [1]
b) If the word "energy" was added to the equation to correctly indicate the energy change in this heat pack reaction, would the word "energy" be placed on the "reactant side" or on the "product side" of the equation? [1]

\[ ____ \text{Fe(s)} + ____ \text{O}_2(g) \rightarrow ____ \text{Fe}_2\text{O}_3(s) \]
Gram formula mass:

1. The gram-formula mass of a compound is 48 grams. The mass of 1.0 mole of this compound is

   1. 1.0 g
   2. 4.8 g
   3. 48 g
   4. 480 g

2. Base your answer on the information below.
   A 150.-gram liquid sample of stearic acid, C_{17}H_{35}COOH, is cooled at a constant rate.
   c Determine the gram-formula mass of stearic acid. [1]

3. Base your answer on the information below.
   Glycine, NH_{2}CH_{2}COOH, is an organic compound found in proteins. Acetamide, CH_{3}CONH_{2}, is an organic compound that is an excellent solvent. Both glycine and acetamide consist of the same four elements, but the compounds have different functional groups.
   Calculate the gram-formula mass of glycine. Your response must include both a numerical setup and the calculated result. [2]

4. What is the gram-formula mass of (NH_{4})_{2}CO_{3}? Use atomic masses rounded to the nearest whole number. [1]

5. In the space provided or on a separate piece of paper, show a correct numerical setup for calculating the formula mass of glucose, C_{6}H_{12}O_{6}.

6. Base your answer on the information below.
   Rust on an automobile door contains Fe_{2}O_{3}(s). The balanced equation representing one of the reactions between iron in the door of the automobile and oxygen in the atmosphere is given below.
   4Fe(s) + 3O_{2}(g) → 2Fe_{2}O_{3}(s)
   Determine the gram-formula mass of the product of this reaction. [1]

7. Base your answer on the information below.
   Hydrogen peroxide, H_{2}O_{2}, is a water-soluble compound. The concentration of an aqueous hydrogen peroxide solution that is 3% by mass H_{2}O_{2} is used as an antiseptic. When the solution is poured on a small cut in the skin, H_{2}O_{2} reacts according to the balanced equation below.
   2H_{2}O_{2} → 2H_{2}O + O_{2}
   Determine the gram-formula mass of H_{2}O_{2}. [1]

Mole/mass/atom calculations:

1. A 1.0-mole sample of krypton gas has a mass of
   1. 19 g
   2. 36 g
   3. 39 g
   4. 84 g

2. Fe(s) + 2HNO_{3}(aq) --> Fe(NO_{3})_{2}(aq) + H_{2}(g)
   What is the total number of oxygen atoms represented in the formula of the iron compound produced?

3. What is the mass of 4.76 moles of Na_{3}PO_{4} (gram-formula mass = 164 grams/mole)?
4. Ozone gas, O$_3$, can be used to kill adult insects in storage bins for grain without damaging the grain. The ozone is produced from oxygen gas, O$_2$, in portable ozone generators located near the storage bins. The concentrations of ozone used are so low that they do not cause any environmental damage. This use of ozone is safer and more environmentally friendly than a method that used bromomethane, CH$_3$Br. However, bromomethane was more effective than ozone because CH$_3$Br killed immature insects as well as adult insects.

Adapted From: The Sunday Gazette (Schenectady, NY) 3/9/03

Determine the total number of moles of CH$_3$Br in 19 grams of CH$_3$Br (gram-formula mass = 95 grams/mol). [1]

5. An unsaturated solution is made by completely dissolving 20.0 grams of NaNO$_3$ in 100.0 grams of water at 20.0°C.

   a In the space on the answer sheet or on a separate piece of paper, show a correct numerical setup for calculating the number of moles of NaNO$_3$ (gram-formula mass = 85.0 grams per mole) used to make this unsaturated solution. [1]
   
   b Determine the minimum mass of NaNO$_3$ that must be added to this unsaturated solution to make a saturated solution at 20.0°C. [1]

   c Identify one process that can be used to recover the NaNO$_3$ from the unsaturated solution. [1]

6. The decomposition of sodium azide, NaN$_3$(s), is used to inflate airbags. On impact, the NaN$_3$(s) is ignited by an electrical spark, producing N$_2$(g) and Na(s). The N$_2$(g) inflates the airbag.

What is the total number of moles present in a 52.0-gram sample of NaN$_3$(s) (gram-formula mass - 65.0 gram/mole)?

7. A student places a 2.50-gram sample of magnesium metal in a bottle and fits the bottle with a 2-hole stopper as shown in the accompanying diagram. Hydrochloric acid is added to the bottle, causing a reaction. As the reaction proceeds, hydrogen gas travels through the tubing to an inverted bottle filled with water, displacing some of the water in the bottle.

In the space provided or on a separate sheet of paper, show a correct numerical setup for calculating the number of moles of magnesium used in the experiment. [1]

8. Determine the mass of 5.20 moles of C$_6$H$_{12}$ (gram-formula mass = 84.2 grams/mole). [1]
**Empirical and molecular formulas:**

1. The molecular formula of glucose is C\(_6\)H\(_{12}\)O\(_6\). What is the empirical formula of glucose?
   - 1. CHO
   - 2. CH\(_2\)O
   - 3. C\(_6\)H\(_{12}\)O\(_6\)
   - 4. C\(_{12}\)H\(_{24}\)O\(_{12}\)

2. A substance has an empirical formula of CH\(_2\) and a molar mass of 56 grams per mole. The molecular formula for this compound is
   - 1. CH\(_2\)
   - 2. C\(_4\)H\(_6\)
   - 3. C\(_6\)H\(_8\)O
   - 4. C\(_{12}\)H\(_{24}\)O\(_{12}\)

3. Which pair consists of a molecular formula and its corresponding empirical formula?
   - 1. C\(_2\)H\(_2\) and CH\(_2\)CH\(_3\)
   - 2. C\(_4\)H\(_6\) and C\(_3\)H\(_2\)
   - 3. P\(_4\)O\(_{10}\) and P\(_2\)O\(_5\)
   - 4. SO\(_2\) and SO\(_3\)

4. What is the empirical formula for a compound with the molecular formula C\(_6\)H\(_{12}\)Cl\(_2\)O\(_2\)?
   - 1. CHClO
   - 2. CH\(_2\)ClO
   - 3. C\(_3\)H\(_6\)ClO
   - 4. C\(_6\)H\(_{12}\)Cl\(_2\)O\(_2\)

5. Which formula is both a molecular and an empirical formula?
   - 1. C\(_6\)H\(_{12}\)O\(_6\)
   - 2. C\(_3\)H\(_4\)O\(_2\)
   - 3. C\(_3\)H\(_8\)O
   - 4. C\(_4\)H\(_8\)

6. Given two formulas representing the same compound: **Formula A** CH\(_3\) and **Formula B** C\(_2\)H\(_6\)
   Which statement describes these formulas?
   - 1. Formulas A and B are both empirical.
   - 2. Formulas A and B are both molecular.
   - 3. Formula A is empirical, and formula B is molecular.
   - 4. Formula A is molecular, and formula B is empirical.

7. Which list consists of types of chemical formulas?
   - 1. atoms, ions, molecules
   - 2. metals, nonmetals, metalloids
   - 3. empirical, molecular, structural
   - 4. synthesis, decomposition, neutralization

8. A compound has a molar mass of 90. grams per mole and the empirical formula CH\(_2\)O. What is the molecular formula of this compound?
   - 1. CH\(_2\)O
   - 2. C\(_2\)H\(_4\)O\(_2\)
   - 3. C\(_3\)H\(_6\)O\(_3\)
   - 4. C\(_4\)H\(_8\)O\(_4\)

9. A compound has the empirical formula CH\(_2\)O and a gram-formula mass of 60. grams per mole.
   What is the molecular formula of this compound?
   - 1. CH\(_2\)O
   - 2. C\(_2\)H\(_4\)O\(_2\)
   - 3. C\(_3\)H\(_8\)O
   - 4. C\(_4\)H\(_8\)O\(_4\)

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10. Given the accompanying formula representing a hydrocarbon:

The molecular formula and the empirical formula for this hydrocarbon are

\[ \text{C}_5\text{H}_{10} \quad \text{and} \quad \text{CH}_2 \]

1. C\textsubscript{5}H\textsubscript{10} and CH\textsubscript{2}  
2. C\textsubscript{4}H\textsubscript{10} and CH\textsubscript{3}  
3. C\textsubscript{4}H\textsubscript{8} and CH\textsubscript{2}  
4. C\textsubscript{4}H\textsubscript{8} and CH\textsubscript{3}

11. In the space provided or on a separate piece of paper, write the empirical formula for the compound C\textsubscript{6}H\textsubscript{12}O\textsubscript{6}.

**Mole ratios:**

1. Given the reaction:
   What is the mole-to-mole ratio between nitrogen gas and hydrogen gas?

   \[ \text{N}_2(g) + 3 \text{H}_2(g) \rightleftharpoons 2 \text{NH}_3(g) \]
   
   1. 1:2  
   2. 1:3  
   3. 2:2  
   4. 2:3

2. Which substance has a chemical formula with the same ratio of metal ions to nonmetal ions as in potassium sulfide?

   1. sodium oxide  
   2. sodium chloride  
   3. magnesium oxide  
   4. magnesium chloride

3. Given the balanced equation representing a reaction: \( \text{F}_2(g) + \text{H}_2(g) \rightarrow 2\text{HF}(g) \)
   What is the mole ratio of H\textsubscript{2}(g) to HF(g) in this reaction?

   1. 1:1  
   2. 1:2  
   3. 2:1  
   4. 2:3

4. Given the balanced equation representing a reaction: \( \text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g) \)
   What is the total number of moles of O\textsubscript{2}(g) required for the complete combustion of 1.5 moles of C\textsubscript{3}H\textsubscript{8}(g)?

   1. 0.30 mol  
   2. 1.5 mol  
   3. 4.5 mol  
   4. 7.5 mol

5. Note: This question may require the use of the *Reference Tables for Physical Setting/Chemistry.*

A sample of a compound contains 65.4 grams of zinc, 12.0 grams of carbon, and 48.0 grams of oxygen. What is the mole ratio of zinc to carbon to oxygen in this compound?

   1. 1:1:2  
   2. 1:1:3  
   3. 1:4:6  
   4. 5:1:4
6. Given the balanced equation representing a reaction: 2CO(g) + O₂(g) → 2CO₂(g)
What is the mole ratio of CO(g) to CO₂(g) in this reaction?
   1. 1:1
   2. 1:2
   3. 2:1
   4. 3:2

7. Base your answer on the balanced chemical equation: 2 H₂O --> 2 H₂ + O₂
What is the total number of moles of O₂ produced when 8 moles of H₂O are completely consumed?

8. Base your answer on the information and equation below.
Antacids can be used to neutralize excess stomach acid. Brand A antacid contains the acid-neutralizing agent magnesium hydroxide, Mg(OH)₂. It reacts with HCl(aq) in the stomach, according to the following balanced equation: 2 HCl(aq) + Mg(OH)₂(s) --> MgCl₂(aq) + 2 H₂O(l)
If a person produces 0.050 mole of excess HCl in the stomach, how many moles of Mg(OH)₂ are needed to neutralize this excess hydrochloric acid? [1]

9. Given the balanced equation: 4Al(s) + 3O₂(g) --> 2Al₂O₃(s)
What is the total number of moles of O₂(g) that must react completely with 8.0 moles of Al(s) in order to form Al₂O₃(s)?

10. Given the equation: 2 H₂(g) + O₂(g) --> 2 H₂O(g)
If 8.0 moles of O₂ are completely consumed, what is the total number of moles of H₂O produced? [1]

11. Given the balanced equation representing a reaction: 2C₂H₆ + 7O₂ → 4CO₂ + 6H₂O
Determine the total number of moles of oxygen that react completely with 8.0 moles of C₂H₆. [1]

Physical behavior of matter / phases:

1. A 1.0-gram sample of which element will uniformly fill a closed 2.0-liter container at STP?
   1. antimony
   2. sulfur
   3. tellurium
   4. xenon

2. At STP, which 2.0-gram sample of matter uniformly fills a 340-milliliter closed container?
   1. Br₂(l)
   2. Fe(NO₃)₂(s)
   3. KCl(aq)
   4. Xe(g)

3. Particles are arranged in a crystal structure in a sample of
   1. H₂(g)
   2. Br₂(l)
   3. Ar(g)
   4. Ag(s)

Vapor pressure / Table H:

1. Which liquid has the lowest vapor pressure at 65°C?
   1. ethanoic acid
   2. ethanol
   3. propanone
   4. water
2. Based on Reference Table H, which sample has the highest vapor pressure?

1. water at 20°C
2. water at 80°C
3. ethanol at 50°C
4. ethanol at 65°C

3. Which sample of water has the lowest vapor pressure?

1. 100 mL at 50°C
2. 200 mL at 30°C
3. 300 mL at 40°C
4. 400 mL at 20°C

4. According to Reference Table H, what is the vapor pressure of propanone at 45°C?

1. 22 kPa
2. 33 kPa
3. 70 kPa
4. 98 kPa

5. Using your knowledge of chemistry and the information in Reference Table H, which statement concerning propanone and water at 50°C is true?

1. Propanone has a higher vapor pressure and stronger intermolecular forces than water.
2. Propanone has a higher vapor pressure and weaker intermolecular forces than water.
3. Propanone has a lower vapor pressure and stronger intermolecular forces than water.
4. Propanone has a lower vapor pressure and weaker intermolecular forces than water.

6. Which substance has the lowest vapor pressure at 75°C?

1. water
2. ethanoic acid
3. propanone
4. ethanol

7. The boiling point of a liquid is the temperature at which the vapor pressure of the liquid is equal to the pressure on the surface of the liquid. What is the boiling point of propanone if the pressure on its surface is 48 kilopascals?

1. 25°C
2. 30°C
3. 35°C
4. 40°C

8. Standard pressure is equal to

1. 1 atm
2. 1 kPa
3. 273 atm
4. 273 kPa

9. At 65°C, which compound has a vapor pressure of 58 kilopascals?

1. ethanoic acid
2. ethanol
3. propanone
4. water

10. Which compound has the lowest vapor pressure at 50°C?

1. ethanoic acid
2. ethanol
3. propanone
4. water
11. Which liquid has the highest vapor pressure at 75°C?
   1. ethanoic acid       3. propanone
   2. ethanol            4. water

12. Based on Reference Table $H$, which substance has the weakest intermolecular forces?
   1. ethanoic acid       3. propanone
   2. ethanol            4. water

13. Base your answer on the article below and on your knowledge of chemistry.

   **Fizzies -- A Splash from the Past**
   They're baaack . . . a splash from the past! Fizzies instant sparkling drink tablets, popular in the 1950s and 1960s, are now back on the market. What sets them apart from other powdered drinks is that they bubble and fizz when placed in water, forming an instant carbonated beverage.
   The fizz in Fizzies is caused by bubbles of carbon dioxide (CO$_2$) gas that are released when the tablet is dropped into water. Careful observation reveals that these bubbles rise to the surface because CO$_2$ gas is much less dense than water. However, not all of the CO$_2$ gas rises to the surface; some of it dissolves in the water. The dissolved CO$_2$ can react with water to form carbonic acid, H$_2$CO$_3$.
   The pH of the Fizzies drink registers between 5 and 6, showing that the resulting solution is clearly acidic. Carbonic acid is found in other carbonated beverages as well. One of the ingredients on any soft drink label is carbonated water, which is another name for carbonic acid. However, in the production of soft drinks, the CO$_2$ is pumped into the solution under high pressure at the bottling plant.

   -- Brian Rohrig
   Excerpted from "Fizzies--A Splash from the Past,"
   Chem Matters, February 1998

CO$_2$ is pumped into the soft drink solution under high pressure. Why is high pressure necessary? [1]

**Intermolecular forces**

1. Organic compounds that are essentially nonpolar and exhibit weak intermolecular forces have
   1. low vapor pressure       3. high boiling points
   2. low melting points       4. high electrical conductivity in solution

2. The strongest forces of attraction occur between molecules of
   1. HCl                      3. HBr
   2. HF                       4. HI

3. At STP, fluorine is a gas and bromine is a liquid because, compared to fluorine, bromine has
   1. stronger covalent bonds  3. weaker covalent bonds
   2. stronger intermolecular forces  4. weaker intermolecular forces

4. The accompanying table shows the normal boiling point of four compounds. Which compound has the strongest intermolecular forces?
5. Which of these substances has the strongest intermolecular forces?

1. H₂O  
2. H₂S  
3. H₂Se  
4. H₂Te

6. Which statement explains why low temperature and high pressure are required to liquefy chlorine gas?

1. Chlorine molecules have weak covalent bonds.  
2. Chlorine molecules have strong covalent bonds.  
3. Chlorine molecules have weak intermolecular forces of attraction.  
4. Chlorine molecules have strong intermolecular forces of attraction.

7. Base your answer on the information below. Naphthalene, a nonpolar substance that sublimes at room temperature, can be used to protect wool clothing from being eaten by moths. Explain, in terms of intermolecular forces, why naphthalene sublimes. [1]

8. Testing of an unknown solid shows that it has the properties listed below.

(1) low melting point  
(2) nearly insoluble in water  
(3) nonconductor of electricity  
(4) relatively soft solid

Explain in terms of attractions between particles why the unknown solid has a low melting point. [1]

9. Explain, in terms of intermolecular forces, why ammonia has a higher boiling point than the other compounds in the table. [1]
10. Crude oil is a mixture of many hydrocarbons that have different numbers of carbon atoms. The use of a fractionating tower allows the separation of this mixture based on the boiling points of the hydrocarbons. To begin the separation process, the crude oil is heated to about 400°C in a furnace, causing many of the hydrocarbons of the crude oil to vaporize. The vaporized mixture is pumped into a fractionating tower that is usually more than 30 meters tall. The temperature of the tower is highest at the bottom. As vaporized samples of hydrocarbons travel up the tower, they cool and condense. The liquid hydrocarbons are collected on trays and removed from the tower. The accompanying diagram illustrates the fractional distillation of the crude oil and the temperature ranges in which the different hydrocarbons condense. Describe the relationship between the strength of the intermolecular forces and the number of carbon atoms in the different hydrocarbon molecules.

11. Carbon and oxygen are examples of elements that exist in more than one form in the same phase. Graphite and diamond are two crystalline arrangements for carbon. The crystal structure of graphite is organized in layers. The bonds between carbon atoms within each layer of graphite are strong. The bonds between carbon atoms that connect different layers of graphite are weak because the shared electrons in these bonds are loosely held by the carbon atoms. The crystal structure of diamond is a strong network of atoms in which all the shared electrons are strongly held by the carbon atoms. Graphite is an electrical conductor, but diamond is not. At 25°C, graphite has a density of 2.2 g/cm³ and diamond has a density of 3.51 g/cm³. The element oxygen can exist as diatomic molecules, O₂, and as ozone, O₃. At standard pressure the boiling point of ozone is 161 K. Explain, in terms of intermolecular forces, the difference in the boiling points of O₂ and O₃ at standard pressure. Your response must include information about both O₂ and O₃. [1]

12. State, in terms of intermolecular forces, why the boiling point of propane at 1 atmosphere is lower than the boiling point of butane at 1 atmosphere. [1]
Ethanol, \( \text{C}_2\text{H}_5\text{OH} \), is a volatile and flammable liquid with a distinct odor at room temperature. Ethanol is soluble in water. The boiling point of ethanol is 78.2°C at 1 atmosphere. Ethanol can be used as a fuel to produce heat energy, as shown by the balanced equation in the accompanying diagram.

Identify one physical property of ethanol, stated in the passage, that can be explained in terms of chemical bonds and intermolecular forces.

\[
\text{C}_2\text{H}_5\text{OH}(\ell) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\ell) + 1367 \text{ kJ}
\]

**Phase changes:**

1. At which Celsius temperature does lead change from a solid to a liquid?
   1. 874°C
   2. 601°C
   3. 328°C
   4. 0°C

2. Which equation represents a physical change?
   (1) \( \text{H}_2\text{O}(\text{s}) + 6.01 \text{ kJ} \rightarrow \text{H}_2\text{O}(\ell) \)
   (2) \( 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g}) + 483.6 \text{ kJ} \)
   (3) \( \text{H}_2(\text{g}) + \text{I}_2(\text{g}) + 53.0 \text{ kJ} \rightarrow 2\text{HI}(\text{g}) \)
   (4) \( \text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) + 66.4 \text{ kJ} \rightarrow 2\text{NO}_2(\text{g}) \)

3. Which process increases the potential energy of the particles of a sample?
   1. condensation
   2. deposition
   3. solidification
   4. vaporization

4. Which equation represents sublimation?
   1. \( \text{I}_2(\text{s}) \rightarrow \text{I}_2(\ell) \)
   2. \( \text{I}_2(\text{s}) \rightarrow \text{I}_2(\ell) \)
   3. \( \text{I}_2(\ell) \rightarrow \text{I}_2(\text{s}) \)
   4. \( \text{I}_2(\ell) \rightarrow \text{I}_2(\ell) \)

5. Given the following equation: \( \text{H}_2\text{O} \text{ (s)} \leftrightarrow \text{H}_2\text{O} \text{ (l)} \). At which temperature will equilibrium exist when the atmospheric pressure is 1 atm?
   1. 0 K
   2. 100 K
   3. 273 K
   4. 373 K
Heating and cooling curves:

1. As ice melts at standard pressure, its temperature remains at 0°C until it has completely melted. Its potential energy
   1. decreases  
   2. increases

2. The accompanying graph represents the relationship between temperature and time as heat is added to a sample of H₂O. Which statement correctly describes the energy of the particles of the sample during interval BC?

   ![Heating Curve for H₂O](image)

   1. Potential energy decreases and average kinetic energy increases.  
   2. Potential energy increases and average kinetic energy increases.  
   3. Potential energy increases and average kinetic energy remains the same.  
   4. Potential energy remains the same and average kinetic energy increases.

3. Base your answer on the information below.
The accompanying graph shows a compound being cooled at a constant rate starting in the liquid phase at 75°C and ending at 15°C. State what is happening to the average kinetic energy of the particles of the sample between minute 2 and minute 6. [1]

![Temperature Changes Over Time](image)

4. Base your answer on the information below.
The temperature of a sample of a substance is increased from 20.0°C to 160.0°C as the sample absorbs heat at a constant rate of 15 kilojoules per minute at standard pressure. The accompanying graph represents the relationship between temperature and time as the sample is heated. What is the boiling point of this sample? [1]
5. Given the heating curve (see image) where substance X starts as a solid below its melting point and is heated uniformly:
   a) On a separate piece of paper identify the process that takes place during line segment DE of the heating curve. [1]
   b) On separate piece of paper identify a line segment in which the average kinetic energy is increasing. [1]
   c) Describe, in terms of particle behavior or energy, what is happening to substance X during line segment BC. [1]

6. Base your answer on the information below.
Heat is added to a sample of liquid water, starting at 80.°C, until the entire sample is a gas at 120.°C. This process, occurring at standard pressure, is represented by the balanced equation below.
H₂O(l) + heat → H₂O(g)
On the diagram or on a separate piece of paper, complete the heating curve for this physical change.
7. Base your answer on the information below and the accompanying diagram.
A 5.00-gram sample of liquid ammonia is originally at 210. K. The diagram of the partial heating curve represents the vaporization of the sample of ammonia at standard pressure due to the addition of heat. The heat is not added at a constant rate.
Some physical constants for ammonia are shown in the accompanying data table.
Describe what is happening to both the potential energy and the average kinetic energy of the molecules in the ammonia sample during time interval BC. Your response must include both potential energy and average kinetic energy. [1]

![Partial Heating Curve for Ammonia](image)

<table>
<thead>
<tr>
<th>Physical Constants for Ammonia</th>
<th>Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Specific heat capacity of NH₃(l)</td>
<td>4.71 J/g-K</td>
</tr>
<tr>
<td>Heat of fusion</td>
<td>332 J/g</td>
</tr>
<tr>
<td>Heat of vaporization</td>
<td>1370 J/g</td>
</tr>
</tbody>
</table>

**Gas laws, KMT, and ideal gas:**

1. If 4.00 moles of oxygen gas, 3.00 moles of hydrogen gas, and 1.00 mole of nitrogen gas are combined in a closed container at standard pressure, what is the partial pressure exerted by the hydrogen gas?
   1. 1.00 atm
   2. 0.125 atm
   3. 3.00 atm
   4. 0.375 atm

2. Base your answer on the information below.
A weather balloon has a volume of 52.5 liters at a temperature of 295 K. The balloon is released and rises to an altitude where the temperature is 252 K. How does this temperature change affect the gas particle motion? [1]

3. Under which conditions of temperature and pressure does oxygen gas behave least like an ideal gas?
   1. low temperature and low pressure
   2. low temperature and high pressure
   3. high temperature and low pressure
   4. high temperature and high pressure
4. A sample of gas occupies a volume of 50.0 milliliters in a cylinder with a movable piston. The pressure of the sample is 0.90 atmosphere and the temperature is 298 K. What is the volume of the sample at STP?

1. 41 mL
2. 49 mL
3. 51 mL
4. 55 mL

5. According to the kinetic molecular theory, the molecules of an ideal gas

1. have a strong attraction for each other
2. have significant volume
3. move in random, constant, straight-line motion
4. are closely packed in a regular repeating pattern

6. Under which conditions of temperature and pressure would a real gas behave most like an ideal gas?

1. 200. K and 50.0 kPa
2. 200. K and 200.0 kPa
3. 600. K and 50.0 kPa
4. 600. K and 200.0 kPa

7. A real gas behaves least like an ideal gas under the conditions of

1. low temperature and low pressure
2. low temperature and high pressure
3. high temperature and low pressure
4. high temperature and high pressure

8. Which statement describes the particles of an ideal gas?

1. The particles move in well-defined, circular paths.
2. When the particles collide, energy is lost.
3. There are forces of attraction between the particles.
4. The volume of the particles is negligible.

9. A sample of gas confined in a cylinder with a movable piston is kept at constant pressure. The volume of the gas doubles when the temperature of the gas is changed from

1. 400. K to 200. K
2. 200. K to 400. K
3. 400.°C to 200.°C
4. 200.°C to 400.°C

10. According to the kinetic molecular theory, the particles of an ideal gas

1. have no potential energy
2. have strong intermolecular forces
3. are arranged in a regular, repeated geometric pattern
4. are separated by great distances, compared to their size

11. Which temperature change would cause a sample of an ideal gas to double in volume while the pressure is held constant?

1. from 400. K to 200. K
2. from 200. K to 400. K
3. from 400.°C to 200.°C
4. from 200.°C to 400.°C
12. As the pressure on the surface of a liquid decreases, the temperature at which the liquid will boil
   1. decreases
   2. increases
   3. remains the same

13. The vapor pressure of a liquid is 0.92 atm at 60°C. The normal boiling point of the liquid could be
   1. 35°C
   2. 45°C
   3. 55°C
   4. 65°C

14. A sample of a gas is contained in a closed rigid cylinder. According to kinetic molecular theory, what occurs when the gas inside the cylinder is heated?
   1. The number of gas molecules increases.
   2. The number of collisions between gas molecules per unit time decreases.
   3. The average velocity of the gas molecules increases.
   4. The volume of the gas decreases.

15. A sample of gas is held at constant pressure. Increasing the kelvin temperature of this gas sample causes the average kinetic energy of its molecules to
   1. decrease and the volume of the gas sample to decrease
   2. decrease and the volume of the gas sample to increase
   3. increase and the volume of the gas sample to decrease
   4. increase and the volume of the gas sample to increase

16. Under which conditions of temperature and pressure does a sample of neon behave most like an ideal gas?
   1. 100 K and 0.25 atm
   2. 100 K and 25 atm
   3. 400 K and 0.25 atm
   4. 400 K and 25 atm

17. According to the kinetic molecular theory, which statement describes the particles in a sample of an ideal gas?
   1. The force of attraction between the gas particles is strong.
   2. The motion of the gas particles is random and straight-line.
   3. The collisions between the gas particles cannot result in a transfer of energy between the particles.
   4. The separation between the gas particles is smaller than the size of the gas particles themselves.

18. A gas sample is at 25°C and 1.0 atmosphere. Which changes in temperature and pressure will cause this sample to behave more like an ideal gas?
   1. decreased temperature and increased pressure
   2. decreased temperature and decreased pressure
   3. increased temperature and increased pressure
   4. increased temperature and decreased pressure
19. Which graph represents the relationship between pressure and volume for a sample of an ideal gas at constant temperature?

![Graphs](image)

20. A cylinder with a movable piston contains a sample of gas having a volume of 6.0 liters at 293 K and 1.0 atmosphere. What is the volume of the sample after the gas is heated to 303 K, while the pressure is held at 1.0 atmosphere?

1. 9.0 L
2. 6.2 L
3. 5.8 L
4. 4.0 L

21. Which graph best shows the relationship between the pressure of a gas and its average kinetic energy at constant volume?

![Graphs](image)

22. The gas volume in the cylinder is 6.2 milliliters and its pressure is 1.4 atmospheres. The piston is then pushed in until the gas volume is 3.1 milliliters while the temperature remains constant.

a Calculate the pressure, in atmospheres, after the change in volume. Record your answer. [2]
b Sketch the general relationship between the pressure and the volume of an ideal gas at constant temperature. [1]
23. Base your answer on the information below.
A weather balloon has a volume of 52.5 liters at a temperature of 295 K. The balloon is released and rises to an altitude where the temperature is 252 K. The original pressure at 295 K was 100.8 kPa and the pressure at the higher altitude at 252 K is 45.6 kPa. Assume the balloon does not burst. On a separate piece of paper, show a correct numerical setup for calculating the volume of the balloon at the higher altitude. [1]

24. Base your answer on the information below.
A piece of magnesium ribbon is reacted with excess hydrochloric acid to produce aqueous magnesium chloride and hydrogen gas. The volume of the dry hydrogen gas produced is 45.6 milliliters. The temperature of the gas is 293 K, and the pressure is 99.5 kilopascals. Calculate the volume this dry hydrogen gas would occupy at STP. Your response must include both a correct numerical setup and the calculated result. [2]

25. Base your answer on the information below.
In a laboratory, a glass tube is filled with hydrogen gas at a very low pressure. When a scientist applies high voltage between metal electrodes in the tube, light is emitted. The scientist analyzes the light with a spectroscope and observes four distinct spectral lines. The accompanying table gives the color, frequency, and energy for each of the four spectral lines. The unit for frequency is hertz, Hz. Identify one condition not mentioned in the passage, under which hydrogen gas behaves most like an ideal gas. [1]

<table>
<thead>
<tr>
<th>Color</th>
<th>Frequency (×10^{14} Hz)</th>
<th>Energy (×10^{-19} J)</th>
</tr>
</thead>
<tbody>
<tr>
<td>red</td>
<td>4.6</td>
<td>3.0</td>
</tr>
<tr>
<td>blue green</td>
<td>6.2</td>
<td>4.1</td>
</tr>
<tr>
<td>blue</td>
<td>6.8</td>
<td>4.6</td>
</tr>
<tr>
<td>violet</td>
<td>7.3</td>
<td>4.8</td>
</tr>
</tbody>
</table>

26. Base your answers to the question on the properties of propanone.
Explain, in terms of molecular energy, why the vapor pressure of propanone increases when its temperature increases.

27. A sample of oxygen gas in one container has a volume of 20.0 milliliters at 297 K and 101.3 kPa. The entire sample is transferred to another container where the temperature is 283 K and the pressure is 94.6 kPa. In the space provided or on a separate piece of paper, show a correct numerical setup for calculating the new volume of this sample of oxygen gas.

**Number of particles in given volume:**

1. A closed container holds 3.0 moles of CO₂ gas at STP. What is the total number of moles of Ne(g) that can be placed in a container of the same size at STP?
   1. 1.0 mole
   2. 1.5 moles
   3. 3.0 moles
   4. 0.0 moles
2. A sealed flask containing 1.0 mole of \( \text{H}_2(g) \) and a sealed flask containing 2.0 moles of \( \text{He}(g) \) are at the same temperature. The two gases must have equal

1. masses
2. volumes
3. average kinetic energies
4. numbers of molecules

3. Which sample at STP has the same number of molecules as 5 liters of \( \text{NO}_2(g) \) at STP?

1. 5 grams of \( \text{H}_2(g) \)
2. 5 liters of \( \text{CH}_4(g) \)
3. 5 moles of \( \text{O}_2(g) \)
4. \( 5 \times 10^{23} \) molecules of \( \text{CO}_2(g) \)

4. At STP, 1.0 liter of helium contains the same total number of atoms as

1. 1.0 L of \( \text{Ne} \)
2. 2.0 L of \( \text{Kr} \)
3. 0.5 L of \( \text{Rn} \)
4. 1.5 L of \( \text{Ar} \)

5. At the same temperature and pressure, 1.0 liter of \( \text{CO}(g) \) and 1.0 liter of \( \text{CO}_2(g) \) have

1. equal masses and the same number of molecules
2. different masses and a different number of molecules
3. equal volumes and the same number of molecules
4. different volumes and a different number of molecules

6. A sample of oxygen gas is sealed in container \( X \). A sample of hydrogen gas is sealed in container \( Z \). Both samples have the same volume, temperature, and pressure. Which statement is true?

1. Container \( X \) contains more gas molecules than container \( Z \).
2. Container \( X \) contains fewer gas molecules than container \( Z \).
3. Containers \( X \) and \( Z \) both contain the same number of gas molecules.
4. Containers \( X \) and \( Z \) both contain the same mass of gas.

7. The accompanying data table gives the temperature and pressure of four different gas samples, each in a 2-liter container. Which two gas samples contain the same total number of particles?

<table>
<thead>
<tr>
<th>Gas Sample</th>
<th>Temperature (K)</th>
<th>Pressure (atm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td>300.</td>
<td>1.20</td>
</tr>
<tr>
<td>Ne</td>
<td>300.</td>
<td>1.00</td>
</tr>
<tr>
<td>\text{CO}_2</td>
<td>200.</td>
<td>1.20</td>
</tr>
<tr>
<td>\text{CH}_4</td>
<td>300.</td>
<td>1.00</td>
</tr>
</tbody>
</table>

1. \( \text{CH}_4 \) and \( \text{CO}_2 \)
2. \( \text{CH}_4 \) and \( \text{Ne} \)
3. \( \text{He} \) and \( \text{CO}_2 \)
4. \( \text{He} \) and \( \text{Ne} \)

8. At STP, which sample contains the same number of molecules as 11.2 liters of \( \text{CO}_2(g) \) at STP?

1. 5.6 L of \( \text{NO}_2(g) \)
2. 7.5 L of \( \text{H}_2(g) \)
3. 11.2 L of \( \text{N}_2(g) \)
4. 22.4 L of \( \text{CO}(g) \)
9. Which gas sample at STP has the same total number of molecules as 2.0 liters of CO\(_2\)(g) at STP?

1. 5.0 L of CO\(_2\)(g)
2. 2.0 L of Cl\(_2\)(g)
3. 3.0 L of H\(_2\)S(g)
4. 6.0 L of He(g)

10. Base your answer to the question on the information below.
The decomposition of sodium azide, NaN\(_3\)(s), is used to inflate airbags. On impact, the NaN\(_3\)(s) is ignited by an electrical spark, producing N\(_2\)(g) and Na(s). The N\(_2\)(g) inflates the airbag.
An inflated airbag has a volume of \(5.00 \times 10^4\) cm\(^3\) at STP. The density of N\(_2\)(g) at STP is 0.00125 g/cm\(^3\). What is the total number of grams of N\(_2\)(g) in the airbag?

**Separation techniques:**

1. Fractional distillation is a technique used to separate complex mixtures of hydrocarbons based on differences in their
   1. heats of fusion
   2. heats of vaporization
   3. melting points
   4. boiling points

2. A dry mixture of KNO\(_3\) and sand could be separated by
   1. adding water to the mixture and filtering
   2. adding water to the mixture and evaporating
   3. heating the mixture to a high temperature
   4. cooling the mixture to a low temperature

3. Which mixture can be separated by using the equipment shown?

4. Which process would most effectively separate two liquids with different molecular polarities?
   1. filtration
   2. fermentation
   3. distillation
   4. conductivity
5. Which property makes it possible to separate the oxygen and the nitrogen from a sample of liquefied air?
   1. boiling point
   2. conductivity
   3. hardness
   4. electronegativity

6. A beaker contains both alcohol and water. These liquids can be separated by distillation because the liquids have different
   1. boiling points
   2. densities
   3. particle sizes
   4. solubilities

7. Petroleum can be separated by distillation because the hydrocarbons in petroleum are
   1. elements with identical boiling points
   2. elements with different boiling points
   3. compounds with identical boiling points
   4. compounds with different boiling points

8. The laboratory process of distillation does not involve
   1. changing a liquid to vapor
   2. changing a vapor to liquid
   3. liquids with different boiling points
   4. liquids with the same boiling points

9. Natural gas is a mixture that includes butane, ethane, methane, and propane. Differences in boiling points can be used to separate the components of natural gas. The boiling points at standard pressure for these components are listed in the accompanying table. Identify a process used to separate the components of natural gas. [1]

<table>
<thead>
<tr>
<th>Component of Natural Gas</th>
<th>Boiling Point at Standard Pressure (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>butane</td>
<td>-0.5</td>
</tr>
<tr>
<td>ethane</td>
<td>-88.6</td>
</tr>
<tr>
<td>methane</td>
<td>-161.6</td>
</tr>
<tr>
<td>propane</td>
<td>-42.1</td>
</tr>
</tbody>
</table>

10. A student prepared two mixtures, each in a labeled beaker. Enough water at 20.°C was used to make 100 milliliters of each mixture. Describe a procedure to physically remove the water from mixture 1. [1]
11. In a laboratory, a student makes a solution by completely dissolving 80.0 grams of KNO₃(s) in 100.0 grams of hot water. The resulting solution has a temperature of 60.0°C. The room temperature in the laboratory is 22°C. Describe a laboratory procedure that can be used to recover the solid solute from the aqueous solution. [1]

**Solutions in general:**

1. A dilute, aqueous potassium nitrate solution is best classified as a  
   1. homogeneous compound  
   2. heterogeneous compound  
   3. homogeneous mixture  
   4. heterogeneous mixture

2. Which sample of matter can be separated into different substances by physical means?  
   1. LiCl(aq)  
   2. LiCl(s)  
   3. NH₃(g)  
   4. NH₃(l)

3. An aqueous solution of sodium chloride is best classified as a  
   1. homogeneous compound  
   2. heterogeneous compound  
   3. homogeneous mixture  
   4. heterogeneous mixture

4. Which formula represents a mixture?  
   1. C₆H₁₂O₆(l)  
   2. C₆H₁₂O₆(s)  
   3. LiCl(aq)  
   4. LiCl(s)

5. Given the equilibrium equation at 298 K. The equation indicates that KNO₃ has formed a saturated solution. Explain, in terms of equilibrium, why the solution is saturated. [1]

\[ \text{KNO}_3(s) + 34.89 \text{kJ} \xrightleftharpoons{H_2O} \text{K}^+(aq) + \text{NO}_3^-(aq) \]

**Solubility:**

1. Base your answer on the accompanying information.  
   Explain, in terms of molecular polarity, why hydrogen chloride is more soluble than methane in water at 20.0°C and standard pressure. [1]

<table>
<thead>
<tr>
<th>Compound</th>
<th>Boiling Point (°C)</th>
<th>Solubility in 100 Grams of H₂O at 20.0°C (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>ammonia</td>
<td>-33.2</td>
<td>56</td>
</tr>
<tr>
<td>methane</td>
<td>-161.5</td>
<td>0.002</td>
</tr>
<tr>
<td>hydrogen chloride</td>
<td>-84.9</td>
<td>72</td>
</tr>
</tbody>
</table>

2. Base your answer on the article below and on your knowledge of chemistry.  
   **Fizzies -- A Splash from the Past**  
   They're baaack . . . a splash from the past! Fizzies instant sparkling drink tablets, popular in the 1950s and 1960s, are now back on the market. What sets them apart from other powdered drinks is that they bubble and fizz when placed in water, forming an instant carbonated beverage.
The fizz in Fizzies is caused by bubbles of carbon dioxide (CO\(_2\)) gas that are released when the tablet is dropped into water. Careful observation reveals that these bubbles rise to the surface because CO\(_2\) gas is much less dense than water. However, not all of the CO\(_2\) gas rises to the surface; some of it dissolves in the water. The dissolved CO\(_2\) can react with water to form carbonic acid, H\(_2\)CO\(_3\). The pH of the Fizzies drink registers between 5 and 6, showing that the resulting solution is clearly acidic. Carbonic acid is found in other carbonated beverages as well. One of the ingredients on any soft drink label is carbonated water, which is another name for carbonic acid. However, in the production of soft drinks, the CO\(_2\) is pumped into the solution under high pressure at the bottling plant.

-- Brian Rohrig

Excerpted from "Fizzies--A Splash from the Past,"
Chem Matters, February 1998

a. Describe the solubility of CO\(_2\) gas in water. [1]
b. Explain your response in terms of the molecular polarities of CO\(_2\)(g) and H\(_2\)O(l). [1]

\[ \text{H}_2\text{O(ℓ)} + \text{CO}_2(\text{aq}) \rightleftharpoons \text{H}_2\text{CO}_3(\text{aq}) \]

3. The health of fish depends on the amount of oxygen dissolved in the water. A dissolved oxygen (DO) concentration between 6 parts per million and 8 parts per million is best for fish health. A DO concentration greater than 1 part per million is necessary for fish survival. Fish health is also affected by water temperature and concentrations of dissolved ammonia, hydrogen sulfide, chloride compounds, and nitrate compounds. Most freshwater fish thrive in water with a pH between 6.5 and 8.5.

A student's fish tank contains fish, green plants, and 3800 grams of fish-tank water with 2.7 \times 10^{-2} gram of dissolved oxygen. Phenolphthalein tests colorless and bromthymol blue tests blue in samples of the fish-tank water.

State how an increase in the temperature of the fish-tank water affects the solubility of oxygen in the water. [1]

**Table F:**

1. According to Table F, which of these salts is least soluble in water?
   1. LiCl  
   2. RbCl  
   3. FeCl\(_2\)  
   4. PbCl\(_2\)

2. According to Reference Table F, which of these compounds is the least soluble in water?
   1. K\(_2\)CO\(_3\)  
   2. KC\(_3\)H\(_2\)O\(_2\)  
   3. Ca\(_3\)(PO\(_4\))\(_2\)  
   4. Ca(NO\(_3\))\(_2\)

3. Which ion, when combined with chloride ions, Cl\(^-\), forms an insoluble substance in water?
   1. Fe\(^{2+}\)  
   2. Mg\(^{2+}\)  
   3. Pb\(^{2+}\)  
   4. Zn\(^{2+}\)

4. According to Table F, which compound is soluble in water?
   1. barium phosphate  
   2. calcium sulfate  
   3. silver iodide  
   4. sodium perchlorate
5. Which of the following compounds is least soluble in water?
   1. copper (II) chloride
   2. aluminum acetate
   3. iron (III) hydroxide
   4. potassium sulfate

6. At STP, which of these substances is most soluble in H₂O?
   1. CCl₄
   2. CO₂
   3. HCl
   4. N₂

7. Which compound is insoluble in water?
   1. BaSO₄
   2. CaCrO₄
   3. KClO₃
   4. Na₂S

8. Which compound is insoluble in water?
   1. KOH
   2. NH₄Cl
   3. Na₃PO₄
   4. PbSO₄

9. Base your answer on the information below.
   A 1.0-gram strip of zinc is reacted with hydrochloric acid in a test tube. The unbalanced equation below represents the reaction.
   \[ \text{Zn(s)} + \text{HCl(aq)} \rightarrow \text{H}_2(g) + \text{ZnCl}_2(aq) \]
   Explain, using information from Reference Table F, why the symbol (aq) is used to describe the product ZnCl₂. [1]

10. Identify one ion from Table F that can combine with Pb²⁺(aq) to produce an insoluble compound. [1]

11. Base your answer on the information and data table.
   Indigestion may be caused by excess stomach acid (hydrochloric acid). Some products used to treat indigestion contain magnesium hydroxide. The magnesium hydroxide neutralizes some of the stomach acid.
   The amount of acid that can be neutralized by three different brands of antacids is shown in the data table.
   Based on Reference Table F, describe the solubility of magnesium hydroxide in water. [1]

<table>
<thead>
<tr>
<th>Antacid Brand</th>
<th>Mass of Antacid Tablet (g)</th>
<th>Volume of HCl(aq) Neutralized (mL)</th>
</tr>
</thead>
<tbody>
<tr>
<td>X</td>
<td>2.00</td>
<td>25.20</td>
</tr>
<tr>
<td>Y</td>
<td>1.20</td>
<td>18.65</td>
</tr>
<tr>
<td>Z</td>
<td>1.75</td>
<td>22.50</td>
</tr>
</tbody>
</table>
12. Base your answer on the article below, the *Reference Tables for Physical Setting/Chemistry*, and your knowledge of chemistry.

In the 1920s, paint used to inscribe the numbers on watch dials was composed of a luminescent (glow-in-the-dark) mixture. The powdered-paint base was a mixture of radium salts and zinc sulfide. As the paint was mixed, the powdered base became airborne and drifted throughout the workroom causing the contents of the workroom, including the painters' clothes and bodies, to glow in the dark. The paint is luminescent because radiation from the radium salts strikes a scintillator. A scintillator is a material that emits visible light in response to ionizing radiation. In watchdial paint, zinc sulfide acts as the scintillator.

Radium present in the radium salts decomposes spontaneously, emitting alpha particles. These particles can cause damage to the body when they enter human tissue. Alpha particles are especially harmful to the blood, liver, lungs, and spleen because they can alter genetic information in the cells. Radium can be deposited in the bones because it substitutes for calcium.

Based on Reference Table *F*, on a separate sheet of paper describe the solubility of zinc sulfide in water. [1]

13. Based on Reference Table *F*, which salt is the strongest electrolyte?

1. CaCO₃  
2. Na₂SO₄  
3. AgCl  
4. Zn₃(PO₄)₂

14. Based on Reference Table *F*, which of these salts is the best electrolyte?

1. sodium nitrate  
2. magnesium carbonate  
3. silver chloride  
4. barium sulfate

**Table G:**

1. According to Reference Table *G*, how many grams of KNO₃ would be needed to saturate 200 grams of water at 70°C?

1. 43 g  
2. 86 g  
3. 134 g  
4. 268 g

2. According to Reference Table *G*, which of these substances is most soluble at 60°C?

1. NaCl  
2. KCl  
3. KClO₃  
4. NH₄Cl

3. A saturated solution of NaNO₃ is prepared at 60°C using 100 grams of water. As this solution is cooled to 10°C, NaNO₃ precipitates (settles) out of the solution. The resulting solution is saturated. Approximately how many grams of NaNO₃ settled out of the original solution?

1. 46 g  
2. 61 g  
3. 85 g  
4. 126 g

4. Which compound is least soluble in water at 60°C?

1. KClO₃  
2. KNO₃  
3. NaCl  
4. NH₄Cl
5. Which compound becomes less soluble in water as the temperature of the solution is increased?
   1. HCl  
   2. KCl  
   3. NaCl  
   4. NH₄Cl

6. One hundred grams of water is saturated with NH₄Cl at 50°C. According to Table G, if the temperature is lowered to 10°C, what is the total amount of NH₄Cl that will precipitate?
   1. 5.0 g  
   2. 17 g  
   3. 30. g  
   4. 50. g

7. A solution contains 35 grams of KNO₃ dissolved in 100 grams of water at 40°C. How much more KNO₃ would have to be added to make it a saturated solution?
   1. 29 g  
   2. 24 g  
   3. 12 g  
   4. 4 g

8. An unsaturated aqueous solution of NH₃ is at 90.°C in 100. grams of water. According to Reference Table G, how many grams of NH₃ could this unsaturated solution contain?
   1. 5 g  
   2. 10. g  
   3. 15 g  
   4. 20. g

9. An unsaturated solution is formed when 80. grams of a salt is dissolved in 100. grams of water at 40.°C. This salt could be
   1. KCl  
   2. KNO₃  
   3. NaCl  
   4. NaNO₃

10. When 5 grams of KCl are dissolved in 50. grams of water at 25°C, the resulting mixture can be described as
    1. heterogeneous and unsaturated  
    2. heterogeneous and supersaturated  
    3. homogeneous and unsaturated  
    4. homogeneous and supersaturated

11. What is the mass of NH₄Cl that must dissolve in 200. grams of water at 50.°C to make a saturated solution?
    1. 26 g  
    2. 42 g  
    3. 84 g  
    4. 104 g

12. Based on Reference Table G, what is the maximum number of grams of KCl(s) that will dissolve in 200 grams of water at 50°C to produce a saturated solution?
    1. 38 g  
    2. 42 g  
    3. 58 g  
    4. 84 g

13. In a laboratory, a student makes a solution by completely dissolving 80.0 grams of KNO₃(s) in 100.0 grams of hot water. The resulting solution has a temperature of 60.°C. The room temperature in the laboratory is 22°C.
    Classify, in terms of saturation, the type of solution made by the student. [1]
14. Given the accompanying data table showing the solubility of salt X:
   a Which salt on Reference Table G is most likely to be salt X? [1]
   b On the graph shown or one similar, scale and label the y-axis including appropriate units. [1]
   c Plot the data from the data table. Surround each point with a small circle and draw a best-fit curve for the solubility of salt X. [1]
   d Using your graph, predict the solubility of salt X at 50° C. [1]
   e If the pressure on the salt solution was increased, what effect would this pressure change have on the solubility of the salt? [1]

<table>
<thead>
<tr>
<th>Temperature (°C)</th>
<th>Mass of Solute per 100 g of H₂O (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>10</td>
<td>22</td>
</tr>
<tr>
<td>25</td>
<td>40.</td>
</tr>
<tr>
<td>30</td>
<td>48</td>
</tr>
<tr>
<td>60</td>
<td>107</td>
</tr>
<tr>
<td>70</td>
<td>135</td>
</tr>
</tbody>
</table>

15. Base your answer on the data table, which shows the solubility of a solid solute. According to Reference Table G, how many grams of KClO₃ must be dissolved in 100 grams of H₂O at 10° C to produce a saturated solution? [1]
Calculating concentration

1. Which preparation produces a 2.0 M solution of $C_6H_{12}O_6$? [molecular mass = 180.0]
   1. 90.0 g of $C_6H_{12}O_6$ dissolved in 500.0 mL of solution
   2. 90.0 g of $C_6H_{12}O_6$ dissolved in 1000. mL of solution
   3. 180.0 g of $C_6H_{12}O_6$ dissolved in 500.0 mL of solution
   4. 180.0 g of $C_6H_{12}O_6$ dissolved in 1000. mL of solution

2. Compared to a 0.1 M aqueous solution of NaCl, a 0.8 M aqueous solution of NaCl has a
   1. higher boiling point and a higher freezing point
   2. higher boiling point and a lower freezing point
   3. lower boiling point and a higher freezing point
   4. lower boiling point and a lower freezing point

3. What is the concentration of a solution, in parts per million, if 0.02 gram of Na$_3$PO$_4$ is dissolved in 1000 grams of water?
   1. 20 ppm
   2. 2 ppm
   3. 0.2 ppm
   4. 0.02 ppm

4. Molarity is defined as the
   1. moles of solute per kilogram of solvent
   2. moles of solute per liter of solution
   3. mass of a solution
   4. volume of a solvent

5. A 3.0 M HCl(aq) solution contains a total of
   1. 3.0 grams of HCl per liter of water
   2. 3.0 grams of HCl per mole of solution
   3. 3.0 moles of HCl per liter of solution
   4. 3.0 moles of HCl per mole of water

6. What is the total mass of solute in 1000. grams of a solution having a concentration of 5 parts per million?
   1. 0.005 g
   2. 0.05 g
   3. 0.5 g
   4. 5 g

7. Which sample of HCl(aq) contains the greatest number of moles of solute particles?
   1. 1.0 L of 2.0 M HCl(aq)
   2. 2.0 L of 2.0 M HCl(aq)
   3. 3.0 L of 0.50 M HCl(aq)
   4. 4.0 L of 0.50 M HCl(aq)

8. The molarity of an aqueous solution of NaCl is defined as the
   1. grams of NaCl per liter of water
   2. grams of NaCl per liter of solution
   3. moles of NaCl per liter of water
   4. moles of NaCl per liter of solution

9. What is the total number of moles of NaCl(s) needed to make 3.0 liters of a 2.0 M NaCl solution?
   1. 1.0 mol
   2. 0.70 mol
   3. 6.0 mol
   4. 8.0 mol
10. A student wants to prepare a 1.0-liter solution of a specific molarity. The student determines that the mass of the solute needs to be 30. grams. What is the proper procedure to follow?

1. Add 30. g of solute to 1.0 L of solvent.
2. Add 30. g of solute to 970. mL of solvent to make 1.0 L of solution.
3. Add 1000. g of solvent to 30. g of solute.
4. Add enough solvent to 30. g of solute to make 1.0 L of solution.

11. How many total moles of KNO₃ must be dissolved in water to make 1.5 liters of a 2.0 M solution?

1. 0.50 mol
2. 2.0 mol
3. 3.0 mol
4. 1.3 mol

12. Which phrase describes the molarity of a solution?

1. liters of solute per mole of solution
2. liters of solution per mole of solution
3. moles of solute per liter of solution
4. moles of solution per liter of solution

13. Which unit can be used to express the concentration of a solution?

1. L/s
2. J/g
3. ppm
4. kPa

14. Bond energy is the amount of energy required to break a chemical bond. The table gives a formula and the carbon-nitrogen bond energy for selected nitrogen compounds.

A 3.2-gram sample of air contains 0.000 74 gram of hydrogen cyanide. Determine the concentration, in parts per million, of the hydrogen cyanide in this sample. [1]

15. On a separate piece of paper, show a correct numerical setup for determining how many liters of a 1.2 M solution can be prepared with 0.50 mole of C₆H₁₂O₆. [1]

16. The health of fish depends on the amount of oxygen dissolved in the water. A dissolved oxygen (DO) concentration between 6 parts per million and 8 parts per million is best for fish health. A DO concentration greater than 1 part per million is necessary for fish survival.

Fish health is also affected by water temperature and concentrations of dissolved ammonia, hydrogen sulfide, chloride compounds, and nitrate compounds. Most freshwater fish thrive in water with a pH between 6.5 and 8.5.

A student's fish tank contains fish, green plants, and 3800 grams of fish-tank water with 2.7 x 10⁻² gram of dissolved oxygen. Phenolphthalein tests colorless and bromthymol blue tests blue in samples of the fish-tank water.

Determine if the DO concentration in the fish tank is healthy for fish. Your response must include:

- a correct numerical setup to calculate the DO concentration in the water in parts per million [1]
- the calculated result [1]
- a statement using your calculated result that tells why the DO concentration in the water is or is not healthy for fish [1]
**Colligative properties:**

1. At standard pressure when NaCl is added to water, the solution will have a
   1. higher freezing point and a lower boiling point than water
   2. higher freezing point and a higher boiling point than water
   3. lower freezing point and higher boiling point than water
   4. lower freezing point and a lower boiling point than water

2. Which solution has the **lowest** freezing point?
   1. 10. g of KI dissolved in 100. g of water
   2. 20. g of KI dissolved in 200. g of water
   3. 30. g of KI dissolved in 100. g of water
   4. 40. g of KI dissolved in 200. g of water

3. Compared to the freezing point and boiling point of water at 1 atmosphere, a solution of a salt and water at 1 atmosphere has a
   1. lower freezing point and a lower boiling point
   2. lower freezing point and a higher boiling point
   3. higher freezing point and a lower boiling point
   4. higher freezing point and a higher boiling point

4. At standard pressure, which element has a freezing point **below** standard temperature?
   1. In
   2. Ir
   3. Hf
   4. Hg

5. At standard pressure, how do the boiling point and freezing point of NaCl(aq) compare to the boiling point and freezing point of H₂O(l)?
   1. Both the boiling point and the freezing point of NaCl(aq) are lower.
   2. Both the boiling point and the freezing point of NaCl(aq) are higher.
   3. The boiling point of NaCl(aq) is lower, and the freezing point of NaCl(aq) is higher.
   4. The boiling point of NaCl(aq) is higher, and the freezing point of NaCl(aq) is lower.

6. Which aqueous solution of KI freezes at the **lowest** temperature?
   1. 1 mol of KI in 500. g of water
   2. 2 mol of KI in 500. g of water
   3. 1 mol of KI in 1000. g of water
   4. 2 mol of KI in 1000. g of water

7. How do the boiling point and freezing point of a solution of water and calcium chloride at standard pressure compare to the boiling point and freezing point of water at standard pressure?
   1. Both the freezing point and boiling point of the solution are higher.
   2. Both the freezing point and boiling point of the solution are lower.
   3. The freezing point of the solution is higher and the boiling point of the solution is lower.
   4. The freezing point of the solution is lower and the boiling point of the solution is higher.
8. In a laboratory, a student makes a solution by completely dissolving 80.0 grams of KNO₃(s) in 100.0 grams of hot water. The resulting solution has a temperature of 60.0°C. The room temperature in the laboratory is 22°C. Compare the boiling point of the solution at standard pressure to the boiling point of water at standard pressure. [1]

9. Ethanol, C₂H₅OH, is a volatile and flammable liquid with a distinct odor at room temperature. Ethanol is soluble in water. The boiling point of ethanol is 78.2°C at 1 atmosphere. Ethanol can be used as a fuel to produce heat energy, as shown by the balanced equation in the accompanying diagram.
At 1 atmosphere, compare the boiling point of pure ethanol to the boiling point of a solution in which a nonvolatile substance is dissolved in ethanol.

\[ \text{C}_2\text{H}_5\text{OH}(\ell) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\ell) + 1367 \text{ kJ} \]

**Thermochemical calculations:**

1. A 36-gram sample of water has an initial temperature of 22°C. After the sample absorbs 1200 joules of heat energy, the final temperature of the sample is
   1. 8.0°C
   2. 14°C
   3. 30°C
   4. 55°C

2. In which equation does the term "heat" represent heat of fusion?
   (1) \( \text{NaCl}(s) + \text{heat} \rightarrow \text{NaCl}(\ell) \)
   (2) \( \text{NaOH(aq)} + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(\ell)} + \text{heat} \)
   (3) \( \text{H}_2\text{O(\ell)} + \text{heat} \rightarrow \text{H}_2\text{O(g)} \)
   (4) \( \text{H}_2\text{O(\ell)} + \text{HCl(g)} \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{heat} \)

3. How much heat energy must be absorbed to completely melt 35.0 grams of H₂O(s) at 0°C?
   1. 9.54 J
   2. 146 J
   3. 11 700 J
   4. 79 100 J

4. What is the total number of joules released when a 5.00-gram sample of water changes from liquid to solid at 0°C?
   1. 334 J
   2. 1670 J
   3. 2260 J
   4. 11,300 J

5. What amount of heat is required to completely melt a 29.95-gram sample of H₂O(s) at 0°C?
   1. \( 334 \text{ J} \)
   2. \( 2260 \text{ J} \)
   3. \( 1.00 \times 10^3 \text{ J} \)
   4. \( 1.00 \times 10^4 \text{ J} \)
6. What is the minimum amount of heat required to completely melt 20.0 grams of ice at its melting point?

   1. 20.0 J  3. 6680 J
   2. 83.6 J  4. 45 200 J

7. a. On a separate piece of paper, calculate the heat released when 25.0 grams of water freezes at 0°C. Show all work. [1]
   b. Record your answer with an appropriate unit. [1]

8. Base your answer on the information below and the accompanying table.
   A substance is a solid at 15°C. A student heated a sample of the solid substance and recorded the temperature at one-minute intervals in the accompanying data table.

   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?

<table>
<thead>
<tr>
<th>Time (min)</th>
<th>0</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
<th>7</th>
<th>8</th>
<th>9</th>
<th>10</th>
<th>11</th>
<th>12</th>
</tr>
</thead>
<tbody>
<tr>
<td>Temperature (°C)</td>
<td>15</td>
<td>32</td>
<td>46</td>
<td>53</td>
<td>53</td>
<td>53</td>
<td>53</td>
<td>53</td>
<td>53</td>
<td>53</td>
<td>60</td>
<td>65</td>
<td></td>
</tr>
</tbody>
</table>

9. Base your answer on the information below.
   At a pressure of 101.3 kilopascals and a temperature of 373 K, heat is removed from a sample of water vapor, causing the sample to change from the gaseous phase to the liquid phase. This phase change is represented by the equation below. 
   \[ \text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(l) + \text{heat} \].
   Determine the total amount of heat released by 5.00 grams of water vapor during this phase change. [1]

10. Base your answer on the information below.
   Heat is added to a 200.-gram sample of \( \text{H}_2\text{O}(s) \) to melt the sample at 0°C. Then the resulting \( \text{H}_2\text{O}(l) \) is heated to a final temperature of 65°C.
   a Determine the total amount of heat required to completely melt the sample. [1]
   b In the space on the answer sheet, show a numerical setup for calculating the total amount of heat required to raise the temperature of the \( \text{H}_2\text{O}(l) \) from 0°C to its final temperature. [1]
   c Compare the amount of heat required to vaporize a 200.-gram sample of \( \text{H}_2\text{O}(l) \) at its boiling point to the amount of heat required to melt a 200.-gram sample of \( \text{H}_2\text{O}(s) \) at its melting point. [1]

11. Base your answer on the information below and the accompanying diagram.
   A 5.00-gram sample of liquid ammonia is originally at 210. K. The diagram of the partial heating curve represents the vaporization of the sample of ammonia at standard pressure due to the addition of heat. The heat is \textit{not} added at a constant rate.
   Some physical constants for ammonia are shown in the accompanying data table.
   a) Determine the total amount of heat required to vaporize this 5.00-gram sample of ammonia at its boiling point. [1]
   b) In the space provided or on a separate piece of paper, calculate the total heat absorbed by the 5.00-gram sample of ammonia during time interval \( AB \). Your response must include \textit{both} a correct numerical setup and the calculated result. [2]
Ethanol, C\(_2\)H\(_5\)OH, is a volatile and flammable liquid with a distinct odor at room temperature. Ethanol is soluble in water. The boiling point of ethanol is 78.2°C at 1 atmosphere. Ethanol can be used as a fuel to produce heat energy, as shown by the balanced equation in the accompanying diagram.

Determine the total amount of heat produced by the complete combustion of 2.00 moles of ethanol.

\[
C_2H_5OH(\ell) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(\ell) + 1367 \text{ kJ}
\]

**Heat flow / exothermic and endothermic:**

1. A student observed that the temperature of water increased when a salt was dissolved in it. The student should conclude that dissolving the salt caused
   1. formation of an acidic solution
   2. formation of a basic solution
   3. an exothermic reaction
   4. an endothermic reaction

2. Which statement correctly describes an endothermic chemical reaction?
   1. the products have higher potential energy than the reactants, and the \(\Delta H\) is negative
   2. the products have higher potential energy than the reactants, and the \(\Delta H\) is positive
   3. the products have lower potential energy than the reactants, and the \(\Delta H\) is negative
   4. the products have lower potential energy than the reactants, and the \(\Delta H\) is positive

3. Which balanced equation represents an endothermic reaction?
   1. C(s) + O\(_2\)(g) \(\rightarrow\) CO\(_2\)(g)
   2. CH\(_4\)(g) + 2O\(_2\)(g) \(\rightarrow\) CO\(_2\)(g) + 2H\(_2\)O(\(l\))
   3. N\(_2\)(g) + 3H\(_2\)(g) \(\rightarrow\) 2NH\(_3\)(g)
   4. N\(_2\)(g) + O\(_2\)(g) \(\rightarrow\) 2NO(g)

---

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4. A thermometer is in a beaker of water. Which statement best explains why the thermometer reading initially increases when LiBr(s) is dissolved in the water?

1. The entropy of the LiBr(aq) is greater than the entropy of the water.
2. The entropy of the LiBr(aq) is less than the entropy of the water.
3. The dissolving of the LiBr(s) in water is an endothermic process.
4. The dissolving of the LiBr(s) in water is an exothermic process.

5. An iron bar at 325 K is placed in a sample of water. The iron bar gains energy from the water if the temperature of the water is

1. 65 K
2. 45 K
3. 65°C
4. 45°C

6. Which equation represents an exothermic reaction at 298 K?

(1) \( \text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g) \)
(2) \( \text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g) \)
(3) \( \text{KNO}_3(s) \xrightarrow{\Delta H} \text{K}^+(aq) + \text{NO}_3^-(aq) \)
(4) \( \text{NH}_4\text{Cl}(s) \xrightarrow{\Delta H} \text{NH}_4^+(aq) + \text{Cl}^-(aq) \)

7. A person with a body temperature of 37°C holds an ice cube with a temperature of 0°C in a room where the air temperature is 20°C. The direction of heat flow is

1. from the person to the ice, only
2. from the person to the ice and air, and from the air to the ice
3. from the ice to the person, only
4. from the ice to the person and air, and from the air to the person

8. In a laboratory where the air temperature is 22°C, a steel cylinder at 100°C is submerged in a sample of water at 40°C. In this system, heat flows from

1. both the air and the water to the cylinder
2. both the cylinder and the air to the water
3. the air to the water and from the water to the cylinder
4. the cylinder to the water and from the water to the air

9. What is the overall result when \( \text{CH}_4(g) \) burns according to this reaction?

\[
\text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g) + \text{CO}_2(g)
\]

1. Energy is absorbed and \( \Delta H \) is negative.
2. Energy is absorbed and \( \Delta H \) is positive.
3. Energy is released and \( \Delta H \) is negative.
4. Energy is released and \( \Delta H \) is positive.
10. Given the reaction shown:

\[ 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O} (\ell) + 571.6 \text{ kJ} \]

What is the approximate \( \Delta H \) for the formation of 1 mole of \( \text{H}_2\text{O} (\ell) \)?

(1) -285.8 kJ
(2) +285.8 kJ
(3) -571.6 kJ
(4) +571.6 kJ

11. Two samples of gold that have different temperatures are placed in contact with one another. Heat will flow spontaneously from a sample of gold at 60°C to a sample of gold that has a temperature of

1. 50°C
2. 60°C
3. 70°C
4. 80°C

12. Given the balanced equation representing a reaction:

\[ \text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g}) + \text{CO}_2(\text{g}) + \text{heat} \]

Which statement is true about energy in this reaction?

1. The reaction is exothermic because it releases heat.
2. The reaction is exothermic because it absorbs heat.
3. The reaction is endothermic because it releases heat.
4. The reaction is endothermic because it absorbs heat.

13. Which statement describes the transfer of heat energy that occurs when an ice cube is added to an insulated container with 100 milliliters of water at 25°C?

1. Both the ice cube and the water lose heat energy.
2. Both the ice cube and the water gain heat energy.
3. The ice cube gains heat energy and the water loses heat energy.
4. The ice cube loses heat energy and the water gains heat energy.

14. Object \( A \) at 40.0°C and object \( B \) at 80.0°C are placed in contact with each other. Which statement describes the heat flow between the objects?

1. Heat flows from object \( A \) to object \( B \).
2. Heat flows from object \( B \) to object \( A \).
3. Heat flows in both directions between the objects.
4. No heat flow occurs between the objects.

15. Given the balanced equation for dissolving \( \text{NH}_4\text{Cl}(\text{s}) \) in water (see accompanying diagram):

\[ \text{NH}_4\text{Cl}(\text{s}) + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ (\text{aq}) + \text{Cl}^- (\text{aq}) \]

A student is holding a test tube containing 5.0 milliliters of water. When a sample of \( \text{NH}_4\text{Cl}(\text{s}) \) is placed in the test tube, the test tube feels colder to the student's hand. Describe the direction of heat flow between the test tube and the hand.

16. A hot pack contains chemicals that can be activated to produce heat. A cold pack contains chemicals that feel cold when activated.

17. Base your answer on the information below.
A hot pack contains chemicals that can be activated to produce heat. A cold pack contains chemicals that feel cold when activated.
a) Based on energy flow, state the type of chemical change that occurs in a hot pack. [1]
b) A cold pack is placed on an injured leg. Indicate the direction of the flow of energy between the leg and the cold pack. [1]

18. Base your answer on the information below.
Given the reaction at equilibrium (see accompanying diagram):
Explain, in terms of energy, why the forward reaction is exothermic.

\[ 2\text{NO}_2(g) \rightleftharpoons \text{N}_2\text{O}_4(g) + 55.3 \text{ kJ} \]

When a person perspires (sweats), the body loses many sodium ions and potassium ions. The evaporation of sweat cools the skin.
After a strenuous workout, people often quench their thirst with sports drinks that contain NaCl and KCl. A single 250-gram serving of one sports drink contains 0.055 gram of sodium ions.
Describe the transfer of energy between the skin and the surroundings as a person perspires and the sweat evaporates. [1]

20. Base your answer on the information below.
Biodiesel is an alternative fuel for vehicles that use petroleum diesel. Biodiesel is produced by reacting vegetable oil with \( \text{CH}_3\text{OH} \). Methyl palmitate, \( \text{C}_{15}\text{H}_{31}\text{COOCH}_3 \), a compound found in biodiesel, is made from soybean oil. One reaction of methyl palmitate with oxygen is represented by the balanced equation below.

\[ 2\text{C}_{15}\text{H}_{31}\text{COOCH}_3 + 49\text{O}_2 \rightarrow 34\text{CO}_2 + 34\text{H}_2\text{O} + \text{energy} \]
State evidence from the balanced equation that indicates the reaction is exothermic. [1]

21. Base your answer on the reaction represented by the balanced equation below.
\[ 2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l) + 571.6 \text{ kJ} \]
Identify the information in this equation that indicates the reaction is exothermic. [1]

22. Given the equation for the dissolving of sodium chloride in water:
When \( \text{NaCl(s)} \) is added to water in a 250-milliliter beaker, the temperature of the mixture is lower than the original temperature of the water. Describe this observation in terms of heat flow. [1]

\[ \text{NaCl(s)} + \text{H}_2\text{O} \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq) \]

23. Base your answer on the information below.
In a laboratory, a student makes a solution by completely dissolving 80.0 grams of KNO\(_3\) in 100.0 grams of hot water. The resulting solution has a temperature of 60.0°C. The room temperature in the laboratory is 22°C.
Describe the direction of heat flow between the solution made by the student and the air in the laboratory. [1]
Table I:

1. According to Table I, which salt releases energy as it dissolves?
   1. KNO₃
   2. LiBr
   3. NH₄NO₃
   4. NaCl

2. Based on Reference Table I, which change occurs when pellets of solid NaOH are added to water and stirred?
   1. The water temperature increases as chemical energy is converted to heat energy.
   2. The water temperature increases as heat energy is stored as chemical energy.
   3. The water temperature decreases as chemical energy is converted to heat energy.
   4. The water temperature decreases as heat energy is stored as chemical energy.

3. Given the potential energy diagram and equation representing the reaction between substances A and D, according to Table I, substance G could be

   ![Potential Energy Diagram]

   1. HI(g)
   2. H₂O(g)
   3. CO₂(g)
   4. C₂H₆(g)

4. According to Table I, which potential energy diagram best represents the reaction that forms H₂O(l) from its elements?

   ![Potential Energy Diagrams]
5. Which reaction releases the greatest amount of energy per 2 moles of product?

1. \(2\text{CO}(g) + \text{O}_2(g) \rightarrow 2\text{CO}_2(g)\)
2. \(4\text{Al}(s) + 3\text{O}_2(g) \rightarrow 2\text{Al}_2\text{O}_3(s)\)
3. \(2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g)\)
4. \(\text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g)\)

**Kinetic energy and temperature:**

1. Which term is defined as a measure of the average kinetic energy of the particles in a sample?
   1. temperature
   2. pressure
   3. thermal energy
   4. chemical energy

2. In which sample is the average kinetic energy of the particles greatest?

   (1) 10. mL of HCl(aq) at 25°C
   (2) 15 mL of HCl(aq) at 20°C
   (3) 10 mL of H2O(l) at 35°C
   (4) 15 mL of H2O(l) at 30°C

3. Which sample of ethanol has particles with the highest average kinetic energy?

   1. 10.0 mL of ethanol at 25°C
   2. 10.0 mL of ethanol at 55°C
   3. 100.0 mL of ethanol at 35°C
   4. 100.0 mL of ethanol at 45°C

4. The temperature of a sample of matter is a measure of the

   1. average kinetic energy of its particles
   2. average potential energy of its particles
   3. total kinetic energy of its particles
   4. total potential energy of its particles

5. Which term is defined as a measure of the average kinetic energy of the particles in a sample of matter?

   1. activation energy
   2. potential energy
   3. temperature
   4. entropy

6. Which sample has particles with the lowest average kinetic energy?

   1. 1.0 g of I2 at 50.0°C
   2. 2.0 g of I2 at 30.0°C
   3. 7.0 g of I2 at 40.0°C
   4. 9.0 g of I2 at 20.0°C

7. In a gaseous system at equilibrium with its surroundings, as molecules of \(A(g)\) collide with molecules of \(B(g)\) without reacting, the total energy of the gaseous system

   1. decreases
   2. increases
   3. remains the same

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

What is the evidence that the average kinetic energy of the particles of this substance is increasing during the first three minutes?
Potential energy diagrams:

1. Base your answer on the information below.
Propane is a fuel that is sold in rigid, pressurized cylinders. Most of the propane in a cylinder is
liquid, with gas in the space above the liquid level. When propane is released from the cylinder, the
propane leaves the cylinder as a gas. Propane gas is used as a fuel by mixing it with oxygen in the air
and igniting the mixture, as represented by the balanced equation shown.
A small amount of methanethiol, which has a distinct odor, is added to the propane to help
consumers detect a propane leak. In methanethiol, the odor is caused by the thiol functional group (–
SH). Methanethiol, CH₃SH, has a structure that is very similar to the structure of methanol.

a In the box provided or on a separate piece of paper, draw a particle diagram to represent propane in
a pressurized cylinder using the key on your answer sheet. Your response must include at least six
molecules of propane in the gas phase and at least six molecules of propane in the liquid phase.

b On the diagram provided or on a separate piece of paper, draw a potential energy diagram for this
reaction.

c Determine the total amount of energy released when 2.50 moles of propane is completely reacted
with oxygen.

d In the space provided or on a separate piece of paper, draw a structural formula for a molecule of
methanethiol.

\[
C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l) + 2219.2 \text{ kJ}
\]
2. Which arrow represents the activation energy of the forward reaction?

![Graph of Reaction Coordinate with Arrows A, B, C, D]

1. \( A \)  
2. \( B \)  
3. \( C \)  
4. \( D \)

3. Which arrow represents the potential energy of the reactants?

![Graph of Reaction Coordinate with Arrows A, B, C, D]

1. \( A \)  
2. \( B \)  
3. \( C \)  
4. \( D \)

4. A catalyst is added to a system at equilibrium. If the temperature remains constant, the activation energy of the forward reaction

1. decreases  
2. increases  
3. remains the same

5. Changes in activation energy during a chemical reaction are represented by a

1. cooling curve  
2. heating curve  
3. ionization energy diagram  
4. potential energy diagram

6. In a chemical reaction, the difference between the potential energy of the products and the potential energy of the reactants is equal to the

1. activation energy  
2. entropy of the system  
3. heat of fusion  
4. heat of reaction

7. In a chemical reaction, the difference between the potential energy of the products and the potential energy of the reactants is equal to the

1. activation energy  
2. kinetic energy  
3. heat of reaction  
4. rate of reaction
8. Which interval on this diagram represents the difference between the potential energy of the products and the potential energy of the reactants?

![Diagram of potential energy vs. reaction coordinate]

9. Each interval on the axis labeled "Potential Energy (kJ)" represents 40 kilojoules. What is the heat of reaction?

![Diagram of potential energy vs. reaction coordinate]

- 1. -120 kJ
- 2. -40 kJ
- 3. +40 kJ
- 4. +160 kJ

10. The activation energy of a chemical reaction can be *decreased* by the addition of

- 1. a catalyst
- 2. an indicator
- 3. electrical energy
- 4. thermal energy

11. When a spark is applied to a mixture of hydrogen and oxygen, the gases react explosively. Which potential energy diagram best represents the reaction?

- (1)
- (2)
- (3)
- (4)
12. Given the reaction at equilibrium and the potential energy diagram shown and complete the diagram for the forward reaction. Be sure your drawing shows the activation energy and the potential energy of the products.

\[ 2\text{NO}_2(g) + 7\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g) + 4\text{H}_2\text{O}(g) + 1127 \text{ kJ} \]

13. On the set of axes provided, sketch the potential energy diagram for an endothermic chemical reaction that shows the activation energy and the potential energy of the reactants and the potential energy of the products. [2]

14. Base your answer on the information shown, which describes the smelting of iron ore, and on your knowledge of chemistry.
In the smelting of iron ore, Fe$_2$O$_3$ is reduced in a blast furnace at high temperature by a reaction with carbon monoxide. Crushed limestone, CaCO$_3$, is also added to the mixture to remove impurities in the ore. The carbon monoxide is formed by the oxidation of carbon (coke), as shown in the reaction shown:

\[ 2 \text{C} + \text{O}_2 \rightarrow 2 \text{ CO} + \text{energy} \]

Liquid iron flows from the bottom of the blast furnace and is processed into different alloys of iron. Using the set of axes provided, sketch a potential energy diagram for the reaction of carbon and oxygen that produces carbon monoxide. [1]
15. Base your answer to the question on the accompanying potential energy diagram.
a) What is the heat of reaction for the forward reaction?
b) What is the activation energy for the forward reaction with the catalyst?
c) Explain, in terms of the function of a catalyst, why the curves on the potential energy diagram for
the catalyzed and uncatalyzed reactions are different.

16. Base your answer on the reaction represented by the balanced equation below.
$\text{2H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{2H}_2\text{O}(l) + 571.6 \text{ kJ}$
On the axes provided, or, if taken online, on a separate piece of paper, draw a potential energy
diagram for the reaction represented by this equation. [1]
17. On the potential energy diagram, draw an arrow to represent the activation energy of the forward reaction. [1]

[Diagram of potential energy with an arrow indicating activation energy]

18. Base your answer on the information below.
The catalytic converter in an automobile changes harmful gases produced during fuel combustion to less harmful exhaust gases. In the catalytic converter, nitrogen dioxide reacts with carbon monoxide to produce nitrogen and carbon dioxide. In addition, some carbon monoxide reacts with oxygen, producing carbon dioxide in the converter. These reactions are represented by the balanced equations below.

Reaction 1:
$$2\text{NO}_2(g) + 4\text{CO}(g) \rightarrow \text{N}_2(g) + 4\text{CO}_2(g) + 1198.4 \text{ kJ}$$

Reaction 2:
$$2\text{CO}(g) + \text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 566.0 \text{ kJ}$$

The potential energy diagram (see image) represents reaction 1 without a catalyst. On the same diagram or a separate piece of paper, draw a dashed line to indicate how potential energy changes when the reaction is catalyzed in the converter. [1]

[Diagram of potential energy with a dashed line indicating catalysis]

Collision theory and reaction rate:

1. A 5.0-gram sample of zinc and a 50.-milliliter sample of hydrochloric acid are used in a chemical reaction. Which combination of these samples has the fastest reaction rate?
   1. a zinc strip and 1.0 M HCl(aq)
   2. a zinc strip and 3.0 M HCl(aq)
   3. zinc powder and 1.0 M HCl(aq)
   4. zinc powder and 3.0 M HCl(aq)

2. For a given reaction, adding a catalyst increases the rate of the reaction by
   1. providing an alternate reaction pathway that has a higher activation energy
   2. providing an alternate reaction pathway that has a lower activation energy
   3. using the same reaction pathway and increasing the activation energy
   4. using the same reaction pathway and decreasing the activation energy
3. How is a chemical reaction affected by the addition of a catalyst?
   1. The activation energy decreases.
   2. The heat of reaction increases.
   3. The number of collisions between particles decreases.
   4. The potential energy of the reactants increases.

4. At 20°C, a 1.2-gram sample of Mg ribbon reacts rapidly with 10.0 milliliters of 1.0 M HCl(aq). Which change in conditions would have caused the reaction to proceed more slowly?
   1. increasing the initial temperature to 25°C
   2. decreasing the concentration of HCl(aq) to 0.1 M
   3. using 1.2 g of powdered Mg
   4. using 2.4 g of Mg ribbon

5. Given the balanced equation representing a reaction:
   \[2\text{HCl(aq)} + \text{Na}_2\text{S}_2\text{O}_3(aq) \rightarrow \text{S(s)} + \text{H}_2\text{SO}_3(aq) + 2\text{NaCl(aq)}\]
   Decreasing the concentration of Na\(_2\)S\(_2\)O\(_3\)(aq) decreases the rate of reaction because the
   1. activation energy decreases
   2. activation energy increases
   3. frequency of effective collisions decreases
   4. frequency of effective collisions increases

6. A chemical reaction between iron atoms and oxygen molecules can only occur if
   1. the particles are heated
   2. the atmospheric pressure decreases
   3. there is a catalyst present
   4. there are effective collisions between the particles

7. For a given chemical reaction, the addition of a catalyst provides a different reaction pathway that
   1. decreases the reaction rate and has a higher activation energy
   2. decreases the reaction rate and has a lower activation energy
   3. increases the reaction rate and has a higher activation energy
   4. increases the reaction rate and has a lower activation energy

8. During a laboratory activity to investigate reaction rate, a student reacts 1.0-gram samples of solid zinc with 10.0-milliliter samples of HCl(aq). The accompanying table shows information about the variables in five experiments the student performed.
   Which two experiments can be used to investigate the effect of the concentration of HCl(aq) on the reaction rate?

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Description of Zinc Sample</th>
<th>HCl(aq) Concentration (M)</th>
<th>Temperature (K)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>lumps</td>
<td>0.10</td>
<td>270.</td>
</tr>
<tr>
<td>2</td>
<td>powder</td>
<td>0.10</td>
<td>270.</td>
</tr>
<tr>
<td>3</td>
<td>lumps</td>
<td>0.10</td>
<td>290.</td>
</tr>
<tr>
<td>4</td>
<td>lumps</td>
<td>1.0</td>
<td>290.</td>
</tr>
<tr>
<td>5</td>
<td>powder</td>
<td>1.0</td>
<td>280.</td>
</tr>
</tbody>
</table>

   1. 1 and 3
   2. 1 and 5
   3. 4 and 2
   4. 4 and 3
9. Based on the nature of the reactants in each of the equations below, which reaction at 25°C will occur at the fastest rate?

1. \( \text{C(s)} + \text{O}_2(g) \rightarrow \text{CO}_2(g) \)

2. \( \text{NaOH(aq)} + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)} \)

3. \( \text{CH}_3\text{OH(l)} + \text{CH}_3\text{COOH(l)} \rightarrow \text{CH}_3\text{COOCH}_3\text{(aq)} + \text{H}_2\text{O(l)} \)

4. \( \text{CaCO}_3(s) \rightarrow \text{CaO(s)} + \text{CO}_2(g) \)

10. A catalyst works by

1. increasing the potential energy of the reactants

2. increasing the energy released during a reaction

3. decreasing the potential energy of the products

4. decreasing the activation energy required for a reaction

11. In each of the four beakers shown in the accompanying image, a 2.0-centimeter strip of magnesium ribbon reacts with 100 milliliters of HCl(aq) under the conditions shown. In which beaker will the reaction occur at the fastest rate?

1. \( \text{Beaker A} \)

2. \( \text{Beaker B} \)

3. \( \text{Beaker C} \)

4. \( \text{Beaker D} \)

12. Which statement best describes how a catalyst increases the rate of a reaction?

1. The catalyst provides an alternate reaction pathway with a higher activation energy.

2. The catalyst provides an alternate reaction pathway with a lower activation energy.

3. The catalyst provides the same reaction pathway with a higher activation energy.

4. The catalyst provides the same reaction pathway with a lower activation energy.
13. Given the balanced equation representing a reaction:
\[ \text{Zn}(s) + 2\text{HCl}(aq) \rightarrow \text{H}_2(g) + \text{ZnCl}_2(aq) \]
Which set of reaction conditions produces \( \text{H}_2(g) \) at the fastest rate?

1. a 1.0-g lump of \( \text{Zn}(s) \) in 50. mL of 0.5 M \( \text{HCl}(aq) \) at 20.°C
2. a 1.0-g lump of \( \text{Zn}(s) \) in 50. mL of 0.5 M \( \text{HCl}(aq) \) at 30.°C
3. 1.0 g of powdered \( \text{Zn}(s) \) in 50. mL of 1.0 M \( \text{HCl}(aq) \) at 20.°C
4. 1.0 g of powdered \( \text{Zn}(s) \) in 50. mL of 1.0 M \( \text{HCl}(aq) \) at 30.°C

14. Why can an increase in temperature lead to more effective collisions between reactant particles and an increase in the rate of a chemical reaction?

1. The activation energy of the reaction increases.
2. The activation energy of the reaction decreases.
3. The number of molecules with sufficient energy to react increases.
4. The number of molecules with sufficient energy to react decreases.

15. Each of four test tubes contains a different concentration of \( \text{HCl}(aq) \) at 25°C. A 1-gram cube of \( \text{Zn} \) is added to each test tube. In which test tube is the reaction occurring at the fastest rate?

16. As the temperature of a chemical reaction in the gas phase is increased, the rate of the reaction increases because

1. fewer particle collisions occur
2. more effective particle collisions occur
3. the required activation energy increases
4. the concentration of the reactants increases

17. A student wishes to investigate how the reaction rate changes with a change in concentration of \( \text{HCl}(aq) \).
   a) Identify one other variable that might affect the rate and should be held constant during this investigation. [1]
   b) Describe the effect of increasing the concentration of \( \text{HCl}(aq) \) on the reaction rate and justify your response in terms of collision theory. [1]

18. Base your answer on the information below.
Given the balanced equation for an organic reaction between butane and chlorine that takes place at 300.°C and 101.3 kilopascals:
C₄H₁₀ + Cl₂ → C₄H₉Cl + HCl
Explain, in terms of collision theory, why the rate of the reaction would decrease if the temperature of the reaction mixture was lowered to 200°C with pressure remaining unchanged.

19. Explain, in terms of collision theory, why the rate of a chemical reaction increases with an increase in temperature. [1]

20. Base your answer on the information below.
Nitrogen gas, hydrogen gas, and ammonia gas are in equilibrium in a closed container at constant temperature and pressure. The accompanying equation represents this equilibrium. The accompanying graph shows the initial concentration of each gas, the changes that occur as a result of adding H₂(g) to the system, and the final concentrations when equilibrium is reestablished. Explain, in terms of collision theory, why the concentration of H₂(g) begins to decrease immediately after more H₂(g) is added to the system. [1]

21. At room temperature, a reaction occurs when KIO₃(aq) is mixed with NaHSO₃(aq) that contains a small amount of starch. The colorless reaction mixture turns dark blue after a period of time that depends on the concentration of the reactants.
In a laboratory, 12 drops of a 0.02 M NaHSO₃(aq) solution containing starch were placed in each of six test tubes. A different number of drops of 0.02 M KIO₃(aq) and enough water to maintain a constant volume were added to each test tube and the time for the dark-blue color to appear was measured. The data were recorded in the accompanying table.
Identify one factor, other than the concentration of the reactants, that would affect the rate of this reaction. [1]

22. Base your answer on the information below.
A student performed a laboratory activity to observe the reaction between aluminum foil and an aqueous copper(II) chloride solution. The reaction is represented by the accompanying balanced equation.
2Al(s) + 3CuCl₂(aq) → 3Cu(s) + 2AlCl₃(aq) + energy
The procedures and corresponding observations for the activities are given in the accompanying table. Describe one change in the procedure that would cause the reaction to occur at a faster rate. [1]

<table>
<thead>
<tr>
<th>Procedure</th>
<th>Observation</th>
</tr>
</thead>
<tbody>
<tr>
<td>In a beaker, completely dissolve 5.00 g of CuCl₂ in 80.0 mL of H₂O.</td>
<td>• The solution is blue green.</td>
</tr>
<tr>
<td>Cut 1.5 g of Al(s) foil into small pieces. Add all the foil to the mixture in the beaker. Stir the contents for 1 minute.</td>
<td>• The surface of Al(s) foil appears partially black. • The beaker feels warm to the touch.</td>
</tr>
<tr>
<td>Observe the beaker and contents after 10 minutes.</td>
<td>• The liquid in the beaker appears colorless. • A reddish-brown solid is soon at the bottom of the beaker. • Some pieces of Al(s) with a partially black coating remain in the beaker.</td>
</tr>
</tbody>
</table>
23. On a separate sheet of paper explain how a catalyst may increase the rate of a chemical reaction.

24. An investigation was conducted to study the effect of the concentration of a reactant on the total time needed to complete a chemical reaction. Four trials of the same reaction were performed. In each trial the initial concentration of the reactant was different. The time needed for the chemical reaction to be completed was measured. The data for each of the four trials are shown in the accompanying table.

In a different experiment involving the same reaction, it was found that an increase in temperature increased the rate of the reaction. Explain this result in terms of collision theory. [1]

<table>
<thead>
<tr>
<th>Trial</th>
<th>Initial Concentration (M)</th>
<th>Reaction Time (s)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.020</td>
<td>11</td>
</tr>
<tr>
<td>2</td>
<td>0.015</td>
<td>14</td>
</tr>
<tr>
<td>3</td>
<td>0.010</td>
<td>23</td>
</tr>
<tr>
<td>4</td>
<td>0.005</td>
<td>58</td>
</tr>
</tbody>
</table>

25. The equilibrium equation shown is related to the manufacture of a bleaching solution. In this equation, Cl\(^-(aq)\) means that chloride ions are surrounded by water molecules.

Explain, in terms of collision theory, why increasing the concentration of Cl\(_2(g)\) increases the concentration of OCl\(^-(aq)\) in this equilibrium system. [1]

\[ \text{Cl}_2(g) + 2\text{OH}^-(aq) \rightleftharpoons \text{OCl}^-(aq) + \text{Cl}^-(aq) + \text{H}_2\text{O}(\ell) \]

26. State two methods to increase the rate of a chemical reaction and explain, in terms of particle behavior, how each method increases the reaction rate. [2]

27. A 1.0-gram strip of zinc is reacted with hydrochloric acid in a test tube. The unbalanced equation below represents the reaction.

\[ \text{Zn}(s) + \text{HCl}(aq) \rightarrow \text{H}_2(g) + \text{ZnCl}_2(aq) \]

Explain, in terms of collision theory, why using 1.0 gram of powdered zinc, instead of the 1.0-gram strip of zinc, would have increased the rate of the reaction. [1]

28. A student wishes to determine how the rate of reaction of magnesium strips with hydrochloric acid, HCl(aq), varies as a function of temperature of the HCl(aq). Give two additional factors, other than the temperature, that could affect the rate of reaction and must be held constant during the experiment. [2]

Entropy and enthalpy:

1. At a pressure of 101.3 kilopascals and a temperature of 373 K, heat is removed from a sample of water vapor, causing the sample to change from the gaseous phase to the liquid phase. This phase change is represented by the equation below.

\[ \text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(\ell) + \text{heat} \]

Explain, in terms of particle arrangement, why entropy decreases during this phase change. [1]
2. Which reaction has the greatest increase in entropy?

\[ \begin{align*}
(1) & \quad 2\text{H}_2\text{O}(\ell) \rightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \\
(2) & \quad 2\text{H}_2\text{O}(\text{g}) \rightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \\
(3) & \quad \text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\ell) \\
(4) & \quad \text{H}_2\text{O}(\ell) \rightarrow \text{H}_2\text{O}(s)
\end{align*} \]

3. Which sample has the greatest entropy?

1. \( \text{NH}_3(\text{g}) \)  
2. \( \text{NH}_3(\ell) \)  
3. \( \text{NH}_3(s) \)  
4. \( \text{NH}_3(aq) \)

4. Which process is accompanied by a decrease in entropy?

1. boiling of water  
2. condensing of water vapor  
3. subliming of iodine  
4. melting of ice

5. As carbon dioxide sublimes, its entropy

1. decreases  
2. increases  
3. remains the same

6. Which 10-milliliter sample of water has the greatest degree of disorder?

1. \( \text{H}_2\text{O}(\text{g}) \) at 120°C  
2. \( \text{H}_2\text{O}(\ell) \) at 80°C  
3. \( \text{H}_2\text{O}(\ell) \) at 20°C  
4. \( \text{H}_2\text{O}(s) \) at 0°C

7. Systems in nature tend to undergo changes toward

1. lower energy and lower entropy  
2. lower energy and higher entropy  
3. higher energy and lower entropy  
4. higher energy and higher entropy

8. Given the balanced equation shown in the diagram:

Which statement best describes this process?

\[ \text{KNO}_3(s) + 34.89 \text{kJ} \rightarrow \text{K}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) \]

1. It is endothermic and entropy increases.  
2. It is endothermic and entropy decreases.  
3. It is exothermic and entropy increases.  
4. It is exothermic and entropy decreases.

9. Systems in nature tend to undergo changes toward

1. lower energy and less disorder  
2. lower energy and more disorder  
3. higher energy and less disorder  
4. higher energy and more disorder

10. Which 1-mole sample has the least entropy?

\[ \begin{align*}
(1) & \quad \text{Br}_2(s) \text{ at } 266 \text{ K} \\
(2) & \quad \text{Br}_2(\ell) \text{ at } 266 \text{ K} \\
(3) & \quad \text{Br}_2(\ell) \text{ at } 332 \text{ K} \\
(4) & \quad \text{Br}_2(\text{g}) \text{ at } 332 \text{ K}
\end{align*} \]
11. Which sample has the *lowest* entropy?

(1) 1 mole of KNO₃(ℓ)  (3) 1 mole of H₂O(ℓ)
(2) 1 mole of KNO₃(s)  (4) 1 mole of H₂O(g)

12. Even though the process is endothermic, snow can sublime. Which tendency in nature accounts for this phase change?

1. a tendency toward greater entropy
2. a tendency toward greater energy
3. a tendency toward less entropy
4. a tendency toward less energy

13. Note: This question may require the use of the *Reference Tables for Physical Setting/Chemistry*. At STP, a sample of which element has the highest entropy?

1. Na(s)  3. Br₂(l)
2. Hg(l)  4. F₂(g)

14. Which list of the phases of H₂O is arranged in order of increasing entropy?

1. ice, steam, and liquid water
2. ice, liquid water, and steam
3. steam, liquid water, and ice
4. steam, ice, and liquid water

15. The entropy of a sample of H₂O increases as the sample changes from a

1. gas to a liquid
2. gas to a solid
3. liquid to a gas
4. liquid to a solid

16. The entropy of a sample of CO₂ increases as the CO₂ changes from

1. gas to liquid
2. gas to solid
3. liquid to solid
4. solid to gas

17. Which equation shows an increase in entropy?

(1) CO₂(g) → CO₂(s)
(2) CO₂(ℓ) → CO₂(g)
(3) CH₃OH(ℓ) → CH₃OH(s)
(4) CH₃OH(g) → CH₃OH(ℓ)

18. The Solvay process is a multistep industrial process used to produce washing soda, Na₂CO₃(s). In the last step of the Solvay process, NaHCO₃(s) is heated to 300°C, producing washing soda, water, and carbon dioxide. This reaction is represented by the balanced equation below.

2NaHCO₃(s) + heat → Na₂CO₃(s) + H₂O(g) + CO₂(g)

State evidence that indicates the entropy of the products is greater than the entropy of the reactant. [1]

19. The balanced equation shown represents the decomposition of potassium chlorate.

2KClO₃(s) → 2KCl(s) + 3O₂(g)

State why the entropy of the reactant is less than the entropy of the products. [1]
20. Given the equation for the dissolving of sodium chloride in water:
On a separate piece of paper describe what happens to entropy during this dissolving process. [1]

\[ \text{NaCl}(s) \xrightarrow{\text{H}_2\text{O}} \text{Na}^+(aq) + \text{Cl}^-(aq) \]

21. Base your answer on the reaction represented by the balanced equation below.
\[ 2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l) + 571.6 \text{kJ} \]
Explain why the entropy of the system decreases as the reaction proceeds. [1]

22. Given the equation
a Name the type of reaction this equation represents. [1]
b Explain, in terms of particle behavior, why entropy is increasing during this reaction. [1]

\[ \text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g) \]

**Equilibrium and LeChatelier's principle:**

1. What occurs when the temperature is increased in a system at equilibrium at constant pressure?
   1. The rate of the forward reaction increases, 3. The rate of the endothermic reaction and the rate of the reverse reaction decreases. increases.
   2. The rate of the forward reaction decreases, 4. The rate of the exothermic reaction and the rate of the reverse reaction increases. decreases.

2. Ammonia is produced commercially by the Haber reaction: . . . (see image)
The formation of ammonia is favored by

\[ \text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g) + \text{heat} \]
   1. an increase in pressure 3. removal of \text{N}_2(g)
   2. a decrease in pressure 4. removal of \text{H}_2(g)

3. Given the equilibrium reaction at STP: . . . (see image)
Which statement correctly describes this system?

\[ \text{N}_2\text{O}_4(g) \rightleftharpoons 2 \text{NO}_2(g) \]
   1. The forward and reverse reaction rates are equal.
   2. The forward and reverse reaction rates are both increasing.
   3. The concentrations of \text{N}_2\text{O}_4 and \text{NO}_2 are equal.
   4. The concentrations of \text{N}_2\text{O}_4 and \text{NO}_2 are both increasing.

5. Which statement about a system at equilibrium is true?
   1. The forward reaction rate is less than the reverse reaction rate.
   2. The forward reaction rate is greater than the reverse reaction rate.
   3. The forward reaction rate is equal to the reverse reaction rate.
   4. The forward reaction rate stops and the reverse reaction rate continues.
6. Which statement must be true about a chemical system at equilibrium?
   1. The forward and reverse reactions stop.
   2. The concentration of reactants and products are equal.
   3. The rate of the forward reaction is equal to the rate of the reverse reaction.
   4. The number of moles of reactants is equal to the number of moles of product.

7. Which factors must be equal in a reversible chemical reaction at equilibrium?
   1. the activation energies of the forward and reverse reactions
   2. the rates of the forward and reverse reactions
   3. the concentrations of the reactants and products
   4. the potential energies of the reactants and products

8. Given the equation \( \text{C}_2\text{H}_5\text{OH}(\ell) \rightleftharpoons \text{C}_2\text{H}_5\text{OH}(g) \) representing a phase change at equilibrium: Which statement is true?
   1. The forward process proceeds faster than the reverse process.
   2. The reverse process proceeds faster than the forward process.
   3. The forward and reverse processes proceed at the same rate.
   4. The forward and reverse processes both stop.

9. Given the equation representing a reaction at equilibrium:
   Which change causes the equilibrium to shift to the right?
   \[ \text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g) + \text{energy} \]
   1. decreasing the concentration of \( \text{H}_2(g) \)
   2. decreasing the pressure
   3. increasing the concentration of \( \text{N}_2(g) \)
   4. increasing the temperature

10. Given the equation representing a reaction: \( \text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g) \)
     Which statement describes this reaction at equilibrium?
     1. The concentration of \( \text{N}_2\text{O}_4(g) \) must equal the concentration of \( \text{NO}_2(g) \).
     2. The concentration of \( \text{N}_2\text{O}_4(g) \) and the concentration of \( \text{NO}_2(g) \) must be constant.
     3. The rate of the forward reaction is greater than the rate of the reverse reaction.
     4. The rate of the reverse reaction is greater than the rate of the forward reaction.

11. Given the equation representing a reaction at equilibrium:
     \( \text{H}_2(g) + \text{I}_2(g) + \text{heat} \rightleftharpoons 2\text{HI}(g) \)
     Which change favors the reverse reaction?
     1. decreasing the concentration of \( \text{HI}(g) \)
     2. decreasing the temperature
     3. increasing the concentration of \( \text{I}_2(g) \)
     4. increasing the pressure

12. Given the equation representing a phase change at equilibrium: \( \text{H}_2\text{O}(s) \rightleftharpoons \text{H}_2\text{O}(\ell) \)
     Which statement describes this equilibrium?
     1. The \( \text{H}_2\text{O}(s) \) melts faster than the \( \text{H}_2\text{O}(\ell) \) freezes.
     2. The \( \text{H}_2\text{O}(\ell) \) freezes faster than the \( \text{H}_2\text{O}(s) \) melts.
     3. The mass of \( \text{H}_2\text{O}(s) \) must equal the mass of \( \text{H}_2\text{O}(\ell) \).
     4. The mass of \( \text{H}_2\text{O}(\ell) \) and the mass of \( \text{H}_2\text{O}(s) \) each remain constant.
13. Given the equation at equilibrium: \( \text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g) + \text{energy} \)
Which changes occur when the temperature of this system is decreased?

1. The concentration of \( \text{H}_2(g) \) increases and the concentration of \( \text{N}_2(g) \) increases.
2. The concentration of \( \text{H}_2(g) \) decreases and the concentration of \( \text{N}_2(g) \) increases.
3. The concentration of \( \text{H}_2(g) \) decreases and the concentration of \( \text{NH}_3(g) \) decreases.
4. The concentration of \( \text{H}_2(g) \) decreases and the concentration of \( \text{NH}_3(g) \) increases.

14. Given the equation representing a closed system: \( \text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g) \)
Which statement describes this system at equilibrium?

1. The volume of the \( \text{NO}_2(g) \) is greater than the volume of the \( \text{N}_2\text{O}_4(g) \).
2. The volume of the \( \text{NO}_2(g) \) is less than the volume of the \( \text{N}_2\text{O}_4(g) \).
3. The rate of the forward reaction and the rate of the reverse reaction are equal.
4. The rate of the forward reaction and the rate of the reverse reaction are unequal.

15. Which statement correctly describes a chemical reaction at equilibrium?

1. The concentrations of the products and reactants are equal.
2. The concentrations of the products and reactants are constant.
3. The rate of the forward reaction is less than the rate of the reverse reaction.
4. The rate of the forward reaction is greater than the rate of the reverse reaction.

16. The solid and liquid phases of water can exist in a state of equilibrium at 1 atmosphere of pressure and a temperature of

1. 0°C
2. 100°C
3. 273°C
4. 373°C

17. Given the reaction \( \text{AgCl}(s) \xrightarrow{\text{H}_2\text{O}} \text{Ag}^+(aq) + \text{Cl}^-(aq) \)
Once equilibrium is reached, which statement is accurate?

1. The concentration of \( \text{Ag}^+(aq) \) is greater than the concentration of \( \text{Cl}^-(aq) \).
2. The \( \text{AgCl}(s) \) will be completely consumed.
3. The rates of the forward and reverse reactions are equal.
4. The entropy of the forward reaction will continue to decrease.

18. Given the reaction system in a closed container at equilibrium and at a temperature of 298 K: \( \text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g) \) The measurable quantities of the gases at equilibrium must be

1. decreasing
2. increasing
3. equal
4. constant

19. A chemical reaction is at equilibrium. Compared to the rate of the forward reaction, the rate of the reverse reaction is

1. faster and more reactant is produced
2. faster and more product is produced
3. the same and the reaction has stopped
4. the same and the reaction continues in both directions
20. Given the reaction at equilibrium: \[ \text{N}_2(g) + 3 \text{H}_2(g) \rightleftharpoons 2 \text{NH}_3(g) + 91.8 \text{ kJ} \]
What occurs when the concentration of \( \text{H}_2(g) \) is increased?

1. The rate of the forward reaction increases and the concentration of \( \text{N}_2(g) \) decreases.
2. The rate of the forward reaction decreases and the concentration of \( \text{N}_2(g) \) increases.
3. The rate of the forward reaction and the concentration of \( \text{N}_2(g) \) both increase.
4. The rate of the forward reaction and the concentration of \( \text{N}_2(g) \) both decrease.

21. Which statement must be true when solution equilibrium occurs?

1. The solution is at STP.
2. The solution is supersaturated.
3. The concentration of the solution remains constant.
4. The masses of the dissolved solute and the undissolved solute are equal.

22. Which two factors must be equal when a chemical reaction reaches equilibrium?

1. the concentration of the reactants and the concentration of the products
2. the number of reactant particles and the number of product particles
3. the rate of the forward reaction and the rate of the reverse reaction
4. the mass of the reactants and the mass of the products

23. Given the equilibrium at 101.3 kPa: \[ \text{H}_2\text{O} (s) \rightleftharpoons \text{H}_2\text{O}(l) \]
At what temperature does this equilibrium occur?

1. 100 K
2. 273 K
3. 298 K
4. 373 K

24. Given the accompanying diagram that shows carbon dioxide in an equilibrium system at a temperature of 298 K and a pressure of 1 atm: Which changes must increase the solubility of the carbon dioxide?

1. increase pressure and decrease temperature
2. increase pressure and increase temperature
3. decrease pressure and decrease temperature
4. decrease pressure and increase temperature

25. Given the reaction shown at equilibrium: \[ \text{N}_2(g) + 3 \text{H}_2(g) \rightleftharpoons 2 \text{NH}_3(g) + 92.05 \text{ kJ} \]

a State the effect on the number of moles of \( \text{N}_2(g) \) if the temperature of the system is increased. [1]
b State the effect on the number of moles of \( \text{H}_2(g) \) if the pressure on the system is increased. [1]
c State the effect on the number of moles of \( \text{NH}_3(g) \) if a catalyst is introduced into the reaction system. Explain why this occurs. [2]
26. Base your answer on the information and equation below. Human blood contains dissolved carbonic acid, $\text{H}_2\text{CO}_3$, in equilibrium with carbon dioxide and water. The equilibrium system is shown. Explain, using LeChatelier's principle, why decreasing the concentration of CO$_2$ decreases the concentration of H$_2$CO$_3$. [1]

$$\text{H}_2\text{CO}_3(\text{aq}) \rightleftharpoons \text{CO}_2(\text{aq}) + \text{H}_2\text{O}(\ell)$$

27. Base your answer on the information below. Given the reaction at equilibrium $2\text{NO}_2(\text{g}) \rightleftharpoons \text{N}_2\text{O}_4(\text{g}) + 55.3 \text{ kJ}$ Explain, in terms of Le Chatelier's principle, why the equilibrium shifts to the right to relieve the stress when the pressure on the system is increased at constant temperature.

28. Base your answer on the information below. Nitrogen gas, hydrogen gas, and ammonia gas are in equilibrium in a closed container at constant temperature and pressure. The accompanying equation represents this equilibrium. The accompanying graph shows the initial concentration of each gas, the changes that occur as a result of adding H$_2$(g) to the system, and the final concentrations when equilibrium is reestablished. $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$

![Concentration of Reaction Gases Versus Time](image)

a) What information on the graph indicates that the system was initially at equilibrium? [1]

b) Explain, in terms of LeChatelier's principle, why the final concentration of NH$_3$(g) is greater than the initial concentration of NH$_3$(g). [1]

29. The equation for the saturated solution equilibrium of potassium nitrate (KNO$_3$) is shown. a On a separate sheet of paper, diagram the products. Use the key provided. Indicate the exact arrangement of the particles you diagram. [2]

b Compare the rate of dissolving KNO$_3$ with the rate of recrystallization of KNO$_3$ for the saturated solution. [1]
30. Base your answer on the information. Given the equilibrium equation at 298 K:
Describe, in terms of LeChatelier's principle, why an increase in temperature increases the solubility of KNO₃. [1]

\[
\text{KNO}_3(s) + \text{H}_2\text{O} \xrightleftharpoons{\text{H}_2\text{O}} \text{K}^+(aq) + \text{NO}_3^-(aq)
\]

31. Base your answer on the information below.
At 550°C, 1.00 mole of CO₂(g) and 1.00 mole of H₂(g) are placed in a 1.00-liter reaction vessel. The substances react to form CO(g) and H₂O(g). Changes in the concentrations of the reactants and the concentrations of the products are shown in the accompanying graph.
What can be concluded from the graph about the concentrations of the reactants and the concentrations of the products between time \( t_1 \) and time \( t_2 \)? [1]
**Acids, bases and salts:**

1. An Arrhenius acid has
   1. only hydroxide ions in solution
   2. only hydrogen ions in solution

2. Which ion is the only negative ion present in an aqueous solution of an Arrhenius base?
   1. hydride ion
   2. hydrogen ion

3. An Arrhenius base yields which ion as the only negative ion in an aqueous solution?
   1. hydride ion
   2. hydrogen ion

4. According to one acid-base theory, a water molecule acts as an acid when the water molecule
   1. accepts an $\text{H}^+$
   2. accepts an $\text{OH}^-$

5. Which two formulas represent Arrhenius acids?
   1. $\text{CH}_3\text{COOH}$ and $\text{CH}_3\text{CH}_2\text{OH}$
   2. $\text{HCl}_2\text{H}_2\text{O}_2$ and $\text{H}_3\text{PO}_4$

6. According to the Arrhenius theory, an acid is a substance that
   1. changes litmus from red to blue
   2. changes phenolphthalein from colorless to pink

7. Given the equation representing a reaction at equilibrium (see accompanying equation):
   The $\text{H}^+$ acceptor for the forward reaction is
   \[ \text{NH}_3(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq) \]
   
   (1) $\text{H}_2\text{O}(l)$
   (2) $\text{NH}_3(g)$
   (3) $\text{NH}_4^+(aq)$
   (4) $\text{OH}^-(aq)$

8. An aqueous solution of lithium hydroxide contains hydroxide ions as the only negative ion in the solution. Lithium hydroxide is classified as an
   1. aldehyde
   2. alcohol
   3. Arrhenius acid
   4. Arrhenius base
9. One alternate acid-base theory states that an acid is an
   1. $\text{H}^+$ donor
   2. $\text{H}^+$ acceptor
   3. $\text{OH}^-$ donor
   4. $\text{OH}^-$ acceptor

10. Given the equation: $\text{HCl(g)} + \text{H}_2\text{O(l)} \rightarrow X(aq) + \text{Cl}^-(aq)$. Which ion is represented by $X$?
   1. hydroxide
   2. hydronium
   3. hypochlorite
   4. perchlorate

11. Which compound when dissolved in water is an Arrhenius acid?
   1. $\text{CH}_3\text{OH}$
   2. $\text{HCl}$
   3. $\text{NaCl}$
   4. $\text{NaOH}$

12. An acid can be defined as an
   1. $\text{H}^+$ acceptor
   2. $\text{H}^+$ donor
   3. $\text{OH}^-$ acceptor
   4. $\text{OH}^-$ donor

13. Potassium hydroxide is classified as an Arrhenius base because KOH contains
   1. $\text{OH}^-$ ions
   2. $\text{O}^2-$ ions
   3. $\text{K}^+$ ions
   4. $\text{H}^+$ ions

14. According to one acid-base theory, an acid is an
   1. $\text{H}^+$ acceptor
   2. $\text{H}^+$ donor
   3. $\text{OH}^-$ acceptor
   4. $\text{OH}^-$ donor

15. Which compound is an Arrhenius base?
   1. $\text{CH}_3\text{OH}$
   2. $\text{CO}_2$
   3. $\text{LiOH}$
   4. $\text{NO}_2$

16. A hydrogen ion, $\text{H}^+$, in aqueous solution may also be written as
   1. $\text{H}_2\text{O}$
   2. $\text{H}_2\text{O}_2$
   3. $\text{H}_3\text{O}^+$
   4. $\text{OH}^-$

17. The compound NaOH(s) dissolves in water to yield
   1. hydroxide ions as the only negative ions
   2. hydroxide ions as the only positive ions
   3. hydronium ions as the only negative ions
   4. hydronium ions as the only positive ions

18. Which formula represents a hydronium ion?
   1. $\text{H}_3\text{O}^+$
   2. $\text{NH}_4^+$
   3. $\text{OH}^-$
   4. $\text{HCO}_3^-$
19. Which compound is an Arrhenius acid?
   1. \( \text{H}_2\text{SO}_4 \)  
   2. \( \text{KCl} \)  
   3. \( \text{NaOH} \)  
   4. \( \text{NH}_3 \)

20. The Arrhenius theory explains the behavior of
   1. acids and bases  
   2. alcohols and amines  
   3. isomers and isotopes  
   4. metals and nonmetals

21. Given the balanced equation representing a reaction:
   \[ \text{NH}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{NH}_4^+(aq) + \text{OH}^-(aq) \]
   According to one acid-base theory, the \( \text{NH}_3(g) \) molecules act as
   1. an acid because they accept \( \text{H}^+ \) ions  
   2. an acid because they donate \( \text{H}^+ \) ions  
   3. a base because they accept \( \text{H}^+ \) ions  
   4. a base because they donate \( \text{H}^+ \) ions

22. One acid-base theory defines a base as an
   1. \( \text{H}^+ \) donor  
   2. \( \text{H}^+ \) acceptor  
   3. \( \text{H} \) donor  
   4. \( \text{H} \) acceptor

23. Which compound is an Arrhenius acid?
   1. \( \text{CaO} \)  
   2. \( \text{HCl} \)  
   3. \( \text{K}_2\text{O} \)  
   4. \( \text{NH}_3 \)

24. According to one acid-base theory, water acts as an acid when an \( \text{H}_2\text{O} \) molecule
   1. accepts an \( \text{H}^+ \)  
   2. donates an \( \text{H}^+ \)  
   3. accepts an \( \text{H}^- \)  
   4. donates an \( \text{H}^- \)

25. Which substance is an Arrhenius acid?
   1. \( \text{NH}_3 \)  
   2. \( \text{KOH} \)  
   3. \( \text{HC}_2\text{H}_3\text{O}_2 \)  
   4. \( \text{CH}_3\text{OH} \)

26. Which substance yields hydroxide ion as the only negative ion in aqueous solution?
   1. \( \text{Mg(OH)}_2 \)  
   2. \( \text{C}_2\text{H}_4\text{(OH)}_2 \)  
   3. \( \text{MgCl}_2 \)  
   4. \( \text{CH}_3\text{Cl} \)

27. A truck carrying concentrated nitric acid overturns and spills its contents. The acid drains into a nearby pond. The pH of the pond water was 8.0 before the spill. After the spill, the pond water is 1,000 times more acidic. Name an ion in the pond water that has increased in concentration due to this spill. [1]

28. A soft-drink bottling plant makes a colorless, slightly acidic carbonated beverage called soda water. During production of the beverage, \( \text{CO}_2(g) \) is dissolved in water at a pressure greater than 1 atmosphere. The bottle containing the solution is capped to maintain that pressure above the solution. As soon as the bottle is opened, fizzing occurs due to \( \text{CO}_2(g) \) being released from the solution. Write the chemical name of the acid in soda water. [1]
29. Base your answer on the information below. 
In liquid water, an equilibrium exists between \( \text{H}_2\text{O}(l) \) molecules, \( \text{H}^+(aq) \) ions, and \( \text{OH}^-(aq) \) ions. A person experiencing acid indigestion after drinking tomato juice can ingest milk of magnesia to reduce the acidity of the stomach contents. Tomato juice has a pH value of 4. Milk of magnesia, a mixture of magnesium hydroxide and water, has a pH value of 10. Identify the negative ion found in milk of magnesia. [1]

30. Base your answer on the article below and on your knowledge of chemistry. 

**Fizzies -- A Splash from the Past**

They're baaack . . . a splash from the past! Fizzes instant sparkling drink tablets, popular in the 1950s and 1960s, are now back on the market. What sets them apart from other powdered drinks is that they bubble and fizz when placed in water, forming an instant carbonated beverage. The fizz in Fizzes is caused by bubbles of carbon dioxide (\( \text{CO}_2 \)) gas that are released when the tablet is dropped into water. Careful observation reveals that these bubbles rise to the surface because \( \text{CO}_2 \) gas is much less dense than water. However, not all of the \( \text{CO}_2 \) gas rises to the surface; some of it dissolves in the water. The dissolved \( \text{CO}_2 \) can react with water to form carbonic acid, \( \text{H}_2\text{CO}_3 \). The pH of the Fizzes drink registers between 5 and 6, showing that the resulting solution is clearly acidic. Carbonic acid is found in other carbonated beverages as well. One of the ingredients on any soft drink label is carbonated water, which is another name for carbonic acid. However, in the production of soft drinks, the \( \text{CO}_2 \) is pumped into the solution under high pressure at the bottling plant.

-- Brian Rohrig

Excerpted from "Fizzies--A Splash from the Past,"

*Chem Matters, February 1998*

What is the only positive ion in an aqueous solution of carbonic acid? [1]

\[
\text{H}_2\text{O}(l) + \text{CO}_2(aq) \rightleftharpoons \text{H}_2\text{CO}_3(aq)
\]

**Electrolytes and nonelectrolytes:**

1. Given the equation for the dissolving of sodium chloride in water: 
   On the piece of paper explain, in terms of particles, why \( \text{NaCl}(s) \) does not conduct electricity. [1]

\[
\text{NaCl}(s) \overset{\text{H}_2\text{O}}{\rightarrow} \text{Na}^+(aq) + \text{Cl}^-(aq)
\]

2. An example of a nonelectrolyte is
   1. \( \text{C}_6\text{H}_12\text{O}_6(aq) \)
   2. \( \text{K}_2\text{SO}_4(aq) \)
   3. \( \text{NaCl}(aq) \)
   4. \( \text{HCl}(aq) \)

3. Which 0.1 M solution contains an electrolyte?
   1. \( \text{C}_6\text{H}_12\text{O}_6(aq) \)
   2. \( \text{CH}_3\text{COOH}(aq) \)
   3. \( \text{CH}_3\text{OH}(aq) \)
   4. \( \text{CH}_3\text{OCH}_3(aq) \)

4. Which formula represents an electrolyte?
   1. \( \text{CH}_3\text{OCH}_3 \)
   2. \( \text{CH}_3\text{OH} \)
   3. \( \text{CH}_3\text{COOH} \)
   4. \( \text{C}_3\text{H}_5\text{CHO} \)
5. Which substance is an electrolyte?
   1. CH₃OH  
   2. C₆H₁₂O₆  
   3. H₂O  
   4. KOH

6. Which compound dissolves in water to form an aqueous solution that can conduct an electric current?
   1. CCl₄  
   2. C₂H₅OH  
   3. CH₃COOH  
   4. CH₄

7. Which laboratory test result can be used to determine if KCl(s) is an electrolyte?
   1. pH of KCl(aq)  
   2. pH of KCl(s)  
   3. electrical conductivity of KCl(aq)  
   4. electrical conductivity of KCl(s)

8. Which compounds can be classified as electrolytes?
   1. alcohols  
   2. alkynes  
   3. organic acids  
   4. saturated hydrocarbons

9. Which aqueous solution is the best conductor of an electrical current?
   1. 0.01 M CH₃OH  
   2. 0.01 M KOH  
   3. 0.1 M CH₃OH  
   4. 0.1 M KOH

10. A substance is classified as an electrolyte because
    1. it has a high melting point  
    2. it contains covalent bonds  
    3. its aqueous solution conducts an electric current  
    4. its aqueous solution has a pH value of 7

11. Which substance is an electrolyte?
    1. CCl₄  
    2. C₃H₆  
    3. HCl  
    4. H₂O

12. Base your answer to the question on the information below.
    Element X is a solid metal that reacts with chlorine to form a water-soluble binary compound.
    Explain, in terms of particles, why an aqueous solution of the binary compound conducts an electric current.

13. Base your answer on the information below.
    When a person perspires (sweats), the body loses many sodium ions and potassium ions. The evaporation of sweat cools the skin.
    After a strenuous workout, people often quench their thirst with sports drinks that contain NaCl and KCl. A single 250.-gram serving of one sports drink contains 0.055 gram of sodium ions.
    State why the salts in sports drinks are classified as electrolytes. [1]
14. Four flasks each contain 100 milliliters of aqueous solutions of equal concentrations at 25°C and 1 atm.
   a Which solutions contain electrolytes? [1]
   b Which solution has the lowest pH? [1]
   c What causes some aqueous solutions to have a low pH? [1]
   d Which solution is most likely to react with an Arrhenius acid to form a salt and water? [1]
   e Which solution has the lowest freezing point? Explain your answer. [2]

![Images of KCl, CH₃OH, Ba(OH)₂, and CH₃COOH]

**Neutralization and titration:**

1. How many milliliters of 12.0 M HCl(aq) must be diluted with water to make exactly 500. mL of 3.00 M hydrochloric acid?
   1. 100. mL
   2. 125. mL
   3. 200. mL
   4. 250. mL

2. Given the reaction shown: HCl(aq) + LiOH(aq) → HOH(ℓ) + LiCl(aq)
   The reaction is best described as
   1. neutralization
   2. synthesis
   3. decomposition
   4. oxidation-reduction

3. If 5.0 milliliters of a 0.20 M HCl solution is required to neutralize exactly 10. milliliters of NaOH, what is the concentration of the base?
   1. 0.10 M
   2. 0.20 M
   3. 0.30 M
   4. 0.40 M

4. Which equation represents a neutralization reaction?
   - (1) Na₂CO₃ + CaCl₂ → 2 NaCl + CaCO₃
   - (2) Ni(NO₃)₂ + H₂S → NiS + 2 HNO₃
   - (3) NaCl + AgNO₃ → AgCl + NaNO₃
   - (4) H₂SO₄ + Mg(OH)₂ → MgSO₄ + 2 H₂O
5. Which reactants form the salt CaSO$_4$(s) in a neutralization reaction?
   1. H$_2$S(g) and Ca(ClO$_4$)$_2$(s)  
   2. H$_2$SO$_3$(aq) and Ca(NO$_3$)$_2$(aq)  
   3. H$_2$SO$_4$(aq) and Ca(OH)$_2$(aq)  
   4. SO$_2$(g) and CaO(s)

6. Based on the equation and the titration results, what is the concentration of the H$_2$SO$_4$(aq)?
   \[\text{H}_2\text{SO}_4(\text{aq}) + 2\text{KOH}(\text{aq}) \rightarrow \text{K}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\ell)\]

<table>
<thead>
<tr>
<th>Titration Experiment Results</th>
</tr>
</thead>
<tbody>
<tr>
<td>volume of H$_2$SO$_4$(aq) used</td>
</tr>
<tr>
<td>concentration of H$_2$SO$_4$(aq)</td>
</tr>
<tr>
<td>volume of KOH(aq) used</td>
</tr>
<tr>
<td>concentration of KOH(aq)</td>
</tr>
</tbody>
</table>

   1. 0.12 M  
   2. 0.16 M  
   3. 0.24 M  
   4. 0.96 M

7. Which word equation represents a neutralization reaction?
   1. base + acid $\rightarrow$ salt + water  
   2. base + salt $\rightarrow$ water + acid  
   3. salt + acid $\rightarrow$ base + water  
   4. salt + water $\rightarrow$ acid + base

8. During which process can 10.0 millimeters of a 0.05 M HCl(aq) solution be used to determine the unknown concentration of a given volume of NaOH(aq) solution?
   1. evaporation  
   2. distillation  
   3. filtration  
   4. titration

9. Which compound is produced when HCl(aq) is neutralized by Ca(OH)$_2$(aq)?
   1. CaCl$_2$  
   2. CaH$_2$  
   3. HClO  
   4. HClO$_2$

10. In which laboratory process is a volume of solution of known concentration used to determine the concentration of another solution?
    1. deposition  
    2. distillation  
    3. filtration  
    4. titration

11. Which solution reacts with LiOH(aq) to produce a salt and water?
    1. KCl(aq)  
    2. CaO(aq)  
    3. NaOH(aq)  
    4. H$_2$SO$_4$(aq)

12. Which volume of 2.0 M NaOH(aq) is needed to completely neutralize 24 milliliters of 1.0 M HCl(aq)?
    1. 6.0 mL  
    2. 12 mL  
    3. 24 mL  
    4. 48 mL
13. When 50. milliliters of an HNO₃ solution is exactly neutralized by 150 milliliters of a 0.50 M solution of KOH, what is the concentration of HNO₃?
   1. 1.0 M
   2. 1.5 M
   3. 3.0 M
   4. 0.5 M

14. Which reaction occurs when hydrogen ions react with hydroxide ions to form water?
   1. substitution
   2. saponification
   3. ionization
   4. neutralization

15. Which equation represents a neutralization reaction?
   1. 4Fe(s) + 3O₂(g) --> 2Fe₂O₃(s)
   2. 2H₂(g) + O₂(g) --> 2H₂O(l)
   3. HNO₃(aq) + KOH(aq) --> KNO₃(aq) + H₂O(l)
   4. AgNO₃(aq) + KCl(aq) --> KNO₃(aq) + AgCl(s)

16. In which laboratory process could a student use 0.10 M NaOH(aq) to determine the concentration of an aqueous solution of HBr?
   1. chromatography
   2. decomposition of the solute
   3. evaporation of the solvent
   4. titration

17. What volume of 0.120 M HNO₃(aq) is needed to completely neutralize 150.0 milliliters of 0.100 M NaOH(aq)?
   1. 62.5 mL
   2. 125 mL
   3. 180. mL
   4. 360. mL

18. Which substance is always a product when an Arrhenius acid in an aqueous solution reacts with an Arrhenius base in an aqueous solution?
   1. HBr
   2. H₂O
   3. KBr
   4. KOH

19. A student completes a titration by adding 12.0 milliliters of NaOH(aq) of unknown concentration to 16.0 milliliters of 0.15 M HCl(aq). What is the molar concentration of the NaOH(aq)?
   1. 0.11 M
   2. 0.20 M
   3. 1.1 M
   4. 5.0 M

20. Base your answer on the information below.
Sulfur dioxide, SO₂, is one gas produced when fossil fuels are burned. When this gas reacts with water in the atmosphere, an acid is produced forming acid rain. The pH of the water in a lake changes when acid rain collects in the lake.
Two samples of the same rainwater are tested using two indicators. Methyl orange is yellow in one sample of this rainwater. Litmus is red in the other sample of this rainwater.
Write the formula for one substance that can neutralize the lake water affected by acid rain. [1]
21. In a titration, 15.65 milliliters of a KOH(aq) solution exactly neutralized 10.00 milliliters of a 1.22 M HCl(aq) solution. 
   a) Show a correct numerical setup for calculating the molarity of the KOH(aq) solution. [1]
   b) Complete the equation for the titration reaction by writing the formula of each product. [1]
      \[
      \text{HCl(aq)} + \text{KOH(aq)} \rightarrow \text{________} + \text{________}
      \]

22. Acid rain is a problem in industrialized countries around the world. Oxides of sulfur and nitrogen are formed when various fuels are burned. These oxides dissolve in atmospheric water droplets that fall to earth as acid rain or acid snow.
   While normal rain has a pH between 5.0 and 6.0 due to the presence of dissolved carbon dioxide, acid rain often has a pH of 4.0 or lower. This level of acidity can damage trees and plants, leach minerals from the soil, and cause the death of aquatic animals and plants.
   If the pH of the soil is too low, then quicklime, CaO, can be added to the soil to increase the pH. Quicklime produces calcium hydroxide when it dissolves in water.
   Samples of acid rain are brought to a laboratory for analysis. Several titrations are performed and it is determined that a 20.0-milliliter sample of acid rain is neutralized with 6.50 milliliters of 0.010 M NaOH. What is the molarity of the H\(^+\) ions in the acid rain?

23. A student recorded the following buret readings during a titration of a base with an acid:
   a) On a separate piece of paper, calculate the molarity of the KOH. Show all work. [1]
   b) Record your answer to the correct number of significant figures. [1]

<table>
<thead>
<tr>
<th></th>
<th>Standard 0.100 M HCl</th>
<th>Unknown KOH</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial reading</td>
<td>9.08 mL</td>
<td>0.55 mL</td>
</tr>
<tr>
<td>Final reading</td>
<td>19.09 mL</td>
<td>5.56 mL</td>
</tr>
</tbody>
</table>

24. A student titrates 60.0 mL of HNO\(_3\)(aq) with 0.30 M NaOH(aq). Phenolphthalein is used as the indicator. After adding 42.2 mL of NaOH(aq), a color change remains for 25 seconds, and the student stops the titration.
   Show a correct numerical setup for calculating the molarity of the HNO\(_3\)(aq). [1]

25. In performing a titration, a student adds three drops of phenolphthalein to a flask containing 25.00 milliliters of HCl(aq). Using a buret, the student slowly adds 0.150 M NaOH(aq) to the flask until one drop causes the indicator to turn light pink. The student determines that a total volume of 20.20 milliliters of NaOH(aq) was used in this titration.
   Calculate the molarity of the HCl(aq) used in this titration. Your response must include both a correct numerical setup and the calculated result. [2]

26. In liquid water, an equilibrium exists between H\(_2\)O(l) molecules, H\(^+\) (aq) ions, and OH\(^-\) (aq) ions. A person experiencing acid indigestion after drinking tomato juice can ingest milk of magnesia to reduce the acidity of the stomach contents. Tomato juice has a pH value of 4. Milk of magnesia, a mixture of magnesium hydroxide and water, has a pH value of 10.
   Complete the accompanying equation for the equilibrium that exists in liquid water. [1]
   \[
   \text{________}_2(\ell) \rightleftharpoons \text{________}_2(\text{aq}) + \text{________}(\text{aq})
   \]
27. Base your answer on the information and data table.
Indigestion may be caused by excess stomach acid (hydrochloric acid). Some products used to treat indigestion contain magnesium hydroxide. The magnesium hydroxide neutralizes some of the stomach acid.
The amount of acid that can be neutralized by three different brands of antacids is shown in the data table.
On a separate piece of paper, show a correct numerical setup for calculating the milliliters of HCl(aq) neutralized per gram of antacid tablet for each brand of antacid. [1]

<table>
<thead>
<tr>
<th>Antacid Brand</th>
<th>Mass of Antacid Tablet (g)</th>
<th>Volume of HCl(aq) Neutralized (mL)</th>
</tr>
</thead>
<tbody>
<tr>
<td>X</td>
<td>2.00</td>
<td>25.20</td>
</tr>
<tr>
<td>Y</td>
<td>1.20</td>
<td>18.65</td>
</tr>
<tr>
<td>Z</td>
<td>1.75</td>
<td>22.50</td>
</tr>
</tbody>
</table>

28. Base your answer to the question on the information below.
In a titration, 3.00 M NaOH(aq) was added to an Erlenmeyer flask containing 25.00 milliliters of HCl(aq) and three drops of phenolphthalein until one drop of the NaOH(aq) turned the solution a light-pink color. The following data were collected by a student performing this titration.
Initial NaOH(aq) buret reading: 14.45 milliliters
Final NaOH(aq) buret reading: 32.66 milliliters
a) What is the total volume of NaOH(aq) that was used in this titration?
b) Show a correct numerical setup for calculating the molarity of the HCl(aq).

29. Base your answer to the question on the information below.
A student was studying the pH differences in samples from two Adirondack streams. The student measured a pH of 4 in stream A and a pH of 6 in stream B.
Identify one compound that could be used to neutralize the sample from stream A.

30. Base your answer on the information below.
Using burets, a student titrated a sodium hydroxide solution of unknown concentration with a standard solution of 0.10 M hydrochloric acid. The data are recorded in the accompanying table.
In the space provided or on a separate sheet of paper, show a correct numerical setup for calculating the molarity of the sodium hydroxide solution. [1]

<table>
<thead>
<tr>
<th>Solution</th>
<th>HCl(aq)</th>
<th>NaOH(aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial Buret Reading (mL)</td>
<td>15.50</td>
<td>5.00</td>
</tr>
<tr>
<td>Final Buret Reading (mL)</td>
<td>25.00</td>
<td>8.80</td>
</tr>
</tbody>
</table>

31. Base your answer on the information below.
In one trial of an investigation, 50.0 milliliters of HCl(aq) of an unknown concentration is titrated with 0.10 M NaOH(aq). During the titration, the total volume of NaOH(aq) added and the corresponding pH value of the reaction mixture are measured and recorded in the accompanying table.
a) Write a balanced equation that represents this neutralization reaction. [1]
b) In another trial, 40.0 milliliters of HCl(aq) is completely neutralized by 20.0 milliliters of this 0.10 M NaOH(aq). Calculate the molarity of the titrated acid in this trial. Your response must include both a numerical setup and the calculated result. [2]

**Calculating pH and pH scale:**

1. Which change in pH represents a hundredfold increase in the concentration of hydronium ions in a solution?
   - 1. pH 1 to pH 2
   - 2. pH 1 to pH 3
   - 3. pH 2 to pH 1
   - 4. pH 3 to pH 1

2. Which pH change represents a hundredfold increase in the concentration of H$_3$O$^+$?
   - 1. pH 5 to pH 7
   - 2. pH 13 to pH 14
   - 3. pH 3 to pH 1
   - 4. pH 4 to pH 3

3. Solution $A$ has a pH of 3 and solution $Z$ has a pH of 6. How many times greater is the hydronium ion concentration in solution $A$ than the hydronium ion concentration in solution $Z$?
   - 1. 100
   - 2. 2
   - 3. 3
   - 4. 1000

4. What is the pH of a solution that has a hydronium ion concentration 100 times greater than a solution with a pH of 4?
   - 1. 5
   - 2. 2
   - 3. 3
   - 4. 6

5. A solution with a pH of 2.0 has a hydronium ion concentration ten times greater than a solution with a pH of
   - 1. 1.0
   - 2. 0.20
   - 3. 3.0
   - 4. 20.
6. Base your answer on the information below.
Sulfur dioxide, \( \text{SO}_2 \), is one gas produced when fossil fuels are burned. When this gas reacts with water in the atmosphere, an acid is produced forming acid rain. The pH of the water in a lake changes when acid rain collects in the lake.
Two samples of the same rainwater are tested using two indicators. Methyl orange is yellow in one sample of this rainwater. Litmus is red in the other sample of this rainwater.
Identify a possible pH value for the rainwater that was tested. [1]

7. Base your answer on the information below.
A truck carrying concentrated nitric acid overturns and spills its contents. The acid drains into a nearby pond. The pH of the pond water was 8.0 before the spill. After the spill, the pond water is 1,000 times more acidic.
What is the new pH of the pond water after the spill? [1]

8. Base your answer on the passage below.
Acid rain is a problem in industrialized countries around the world. Oxides of sulfur and nitrogen are formed when various fuels are burned. These oxides dissolve in atmospheric water droplets that fall to earth as acid rain or acid snow.
While normal rain has a pH between 5.0 and 6.0 due to the presence of dissolved carbon dioxide, acid rain often has a pH of 4.0 or lower. This level of acidity can damage trees and plants, leach minerals from the soil, and cause the death of aquatic animals and plants.
If the pH of the soil is too low, then quicklime, \( \text{CaO} \), can be added to the soil to increase the pH. Quicklime produces calcium hydroxide when it dissolves in water.
A sample of wet soil has a pH of 4.0. After the addition of quicklime, the \( \text{H}^+ \) ion concentration of the soil is 1/100 of the original \( \text{H}^+ \) ion concentration of the soil. What is the new pH of the soil sample?

9. Base your answer on the information below.
In liquid water, an equilibrium exists between \( \text{H}_2\text{O}(l) \) molecules, \( \text{H}^+(aq) \) ions, and \( \text{OH}^-(aq) \) ions. A person experiencing acid indigestion after drinking tomato juice can ingest milk of magnesia to reduce the acidity of the stomach contents. Tomato juice has a pH value of 4. Milk of magnesia, a mixture of magnesium hydroxide and water, has a pH value of 10.
Compare the hydrogen ion concentration in tomato juice to the hydrogen ion concentration in milk of magnesia. [1]

10. Base your answer on the information below.
Some carbonated beverages are made by forcing carbon dioxide gas into a beverage solution. When a bottle of one kind of carbonated beverage is first opened, the beverage has a pH value of 3.
a) State, in terms of the pH scale, why this beverage is classified as acidic. [1]
b) After the beverage bottle is left open for several hours, the hydronium ion concentration in the beverage solution decreases to 1/1000 of the original concentration. Determine the new pH of the beverage solution. [1]

11. Base your answer to the question on the information below.
A student was studying the pH differences in samples from two Adirondack streams. The student measured a pH of 4 in stream \( A \) and a pH of 6 in stream \( B \).
Compare the hydronium ion concentration in stream \( A \) to the hydronium ion concentration in stream \( B \).
12. Base your answer on the information below.

Three bottles of liquids labeled 1, 2, and 3 were found in a storeroom. One of the liquids is known to be drain cleaner. Drain cleaners commonly contain KOH or NaOH. The pH of each liquid at 25°C was determined with a pH meter. The accompanying table shows the test results.

Explain how the pH results in this table enable a student to correctly conclude that bottle 3 contains the drain cleaner. [1]

<table>
<thead>
<tr>
<th>Bottle</th>
<th>pH of Liquid</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>3.8</td>
</tr>
<tr>
<td>2</td>
<td>7.0</td>
</tr>
<tr>
<td>3</td>
<td>12.8</td>
</tr>
</tbody>
</table>

**Indicators / Table M:**

1. A student was given four unknown solutions. Each solution was checked for conductivity and tested with phenolphthalein. The results are shown in the data table. Based on the data table, which unknown solution could be 0.1 M NaOH?

<table>
<thead>
<tr>
<th>Solution</th>
<th>Conductivity</th>
<th>Color with Phenolphthalein</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>Good</td>
<td>Colorless</td>
</tr>
<tr>
<td>B</td>
<td>Poor</td>
<td>Colorless</td>
</tr>
<tr>
<td>C</td>
<td>Good</td>
<td>Pink</td>
</tr>
<tr>
<td>D</td>
<td>Poor</td>
<td>Pink</td>
</tr>
</tbody>
</table>

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

2. A student tested a 0.1 M aqueous solution and made the following observations:
   - conducts electricity
   - turns blue litmus to red
   - reacts with Zn(s) to produce gas bubbles

Which compound could be the solute in this solution?

1. CH$_3$OH  3. HBr
2. LiBr        4. LiOH

3. Which indicator would best distinguish between a solution with a pH of 3.5 and a solution with a pH of 5.5?

1. bromthymol blue  3. litmus
2. bromcresol green 4. thymol blue
4. Which statement correctly describes a solution with a pH of 9?
   1. It has a higher concentration of H$_3$O$^+$ than OH$^-$ and causes litmus to turn blue.
   2. It has a higher concentration of OH$^-$ than H$_3$O$^+$ and causes litmus to turn blue.
   3. It has a higher concentration of H$_3$O$^+$ than OH$^-$ and causes methyl orange to turn yellow.
   4. It has a higher concentration of OH$^-$ than H$_3$O$^+$ and causes methyl orange to turn red.

5. The accompanying table shows the color of the indicators methyl orange and litmus in two samples of the same solution. Which pH value is consistent with the indicator results?

<table>
<thead>
<tr>
<th>Indicator</th>
<th>Color Result from the Indicator Test</th>
</tr>
</thead>
<tbody>
<tr>
<td>methyl orange</td>
<td>yellow</td>
</tr>
<tr>
<td>litmus</td>
<td>red</td>
</tr>
</tbody>
</table>

   1. 1
   2. 5
   3. 3
   4. 10

6. Which indicator is blue in a solution that has a pH of 5.6?
   1. brom cresol green
   2. brom thymol blue
   3. methyl orange
   4. thymol blue

7. Based on the results of testing colorless solutions with indicators, which solution is most acidic?
   1. a solution in which brom thymol blue is blue
   2. a solution in which brom cresol green is blue
   3. a solution in which phenolphthalein is pink
   4. a solution in which methyl orange is red

8. A student is given two beakers, each containing an equal amount of clear, odorless liquid. One solution is acidic and the other is basic.
   a State two safe methods of distinguishing the acid solution from the base solution. [2]
   b For each method, state the results of both the testing of the acid solution and the testing of the base solution. [2]

9. Base your answer on the information below.
   A truck carrying concentrated nitric acid overturns and spills its contents. The acid drains into a nearby pond. The pH of the pond water was 8.0 before the spill. After the spill, the pond water is 1,000 times more acidic.
   What color would brom thymol blue be at this new pH? [1]

10. Base your answer on the information below.
    A student titrates 60.0 mL of HNO$_3$(aq) with 0.30 M NaOH(aq). Phenolphthalein is used as the indicator. After adding 42.2 mL of NaOH(aq), a color change remains for 25 seconds, and the student stops the titration.
    What color change does phenolphthalein undergo during this titration? [1]
11. Base your answer on the information below.
In liquid water, an equilibrium exists between $H_2O(l)$ molecules, $H^+(aq)$ ions, and $OH^-(aq)$ ions. A person experiencing acid indigestion after drinking tomato juice can ingest milk of magnesia to reduce the acidity of the stomach contents. Tomato juice has a pH value of 4. Milk of magnesia, a mixture of magnesium hydroxide and water, has a pH value of 10.
What is the color of thymol blue indicator when placed in a sample of milk of magnesia? [1]

12. Base your answer on the information below.
Some carbonated beverages are made by forcing carbon dioxide gas into a beverage solution. When a bottle of one kind of carbonated beverage is first opened, the beverage has a pH value of 3. Using Table $M$, identify one indicator that is yellow in a solution that has the same pH value as this beverage. [1]

13. Base your answer to the question on the information below.
A student was studying the pH differences in samples from two Adirondack streams. The student measured a pH of 4 in stream $A$ and a pH of 6 in stream $B$.
What is the color of bromthymol blue in the sample from stream $A$?

14. Base your answer on the information below.
Three bottles of liquids labeled 1, 2, and 3 were found in a storeroom. One of the liquids is known to be drain cleaner. Drain cleaners commonly contain KOH or NaOH. The pH of each liquid at 25$^\circ$C was determined with a pH meter. The accompanying table shows the test results.
Explain, in terms of the pH values, why thymol blue is not a suitable indicator to distinguish between the contents of bottle 1 and bottle 2. [1]

<table>
<thead>
<tr>
<th>pH Test Results</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bottle</td>
</tr>
<tr>
<td>-------</td>
</tr>
<tr>
<td>1</td>
</tr>
<tr>
<td>2</td>
</tr>
<tr>
<td>3</td>
</tr>
</tbody>
</table>

15. What color is bromcresol green after it is added to a sample of NaOH(aq)? [1]

**Oxidation numbers:**

1. In which substance is the oxidation number of Cl equal to +1?
   1. $Cl_2$ 
   2. $Cl_2O$ 
   3. $AlCl_3$ 
   4. $HClO_2$

2. What is the oxidation state of nitrogen in NaNO$_2$?
   1. +1 
   2. +2 
   3. +3 
   4. +4
3. Given the balanced equation representing a reaction: \( \text{Fe}_2\text{O}_3 + 2\text{Al} \rightarrow \text{Al}_2\text{O}_3 + 2\text{Fe} \)
During this reaction, the oxidation number of Fe changes from
1. +2 to 0 as electrons are transferred
2. +2 to 0 as protons are transferred
3. +3 to 0 as electrons are transferred
4. +3 to 0 as protons are transferred

4. Given the balanced equation representing a reaction: \( 2\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + 3\text{O}_2(g) \)
The oxidation state of chlorine in this reaction changes from
1. -1 to +1
2. -1 to +5
3. +1 to -1
4. +5 to -1

5. What is the oxidation state of nitrogen in the compound \( \text{NH}_4\text{Br} \)?
1. -1
2. +2
3. -3
4. +4

6. What is the oxidation number of chromium in \( \text{K}_2\text{Cr}_2\text{O}_7 \)?
1. +6
2. +2
3. +7
4. +12

7. Human blood contains dissolved carbonic acid, \( \text{H}_2\text{CO}_3 \), in equilibrium with carbon dioxide and water. The equilibrium system is shown.
What is the oxidation number of carbon in \( \text{H}_2\text{CO}_3(aq) \)? [1]
\[
\text{H}_2\text{CO}_3(aq) \rightleftharpoons \text{CO}_2(aq) + \text{H}_2\text{O}(l)
\]

8. Aluminum is one of the most abundant metals in Earth's crust. The aluminum compound found in bauxite ore is \( \text{Al}_2\text{O}_3 \). Over one hundred years ago, it was difficult and expensive to isolate aluminum from bauxite ore. In 1886, a brother and sister team, Charles and Julia Hall, found that molten (melted) cryolite, \( \text{Na}_3\text{AlF}_6 \), would dissolve bauxite ore. Electrolysis of the resulting mixture caused the aluminum ions in the \( \text{Al}_2\text{O}_3 \) to be reduced to molten aluminum metal. This less expensive process is known as the Hall process.
Write the oxidation state for each of the elements in cryolite.
\[
\text{Na}_3\text{AlF}_6 \quad \text{Na:} \quad \text{( } \quad \text{Al:} \quad \text{( } \quad \text{F:} \quad \text{( }
\]

9. Two sources of copper are cuprite, which has the IUPAC name copper(I) oxide, and malachite, which has the formula \( \text{Cu}_2\text{CO}_3(\text{OH})_2 \). Copper is used in home wiring and electric motors because it has good electrical conductivity. Other uses of copper not related to its electrical conductivity include coins, plumbing, roofing, and cooking pans. Aluminum is also used for cooking pans.
At room temperature, the electrical conductivity of a copper wire is 1.6 times greater than an aluminum wire with the same length and cross-sectional area. At room temperature, the heat conductivity of copper is 1.8 times greater than the heat conductivity of aluminum. At STP, the density of copper is 3.3 times greater than the density of aluminum.
Determine the oxidation number of oxygen in the carbonate ion found in malachite. [1]
10. Base your answer on the information below.
The balanced equation shown represents the decomposition of potassium chlorate.
\[ 2\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + 3\text{O}_2(g) \]
Determine the oxidation number of chlorine in the reactant in the equation. [1]

11. Base your answer on the information shown, which describes the smelting of iron ore, and on your knowledge of chemistry.
In the smelting of iron ore, \( \text{Fe}_2\text{O}_3 \) is reduced in a blast furnace at high temperature by a reaction with carbon monoxide. Crushed limestone, \( \text{CaCO}_3 \), is also added to the mixture to remove impurities in the ore. The carbon monoxide is formed by the oxidation of carbon (coke), as shown in the reaction shown:
\[ 2\text{C} + \text{O}_2 \rightarrow 2\text{CO} + \text{energy} \]
Liquid iron flows from the bottom of the blast furnace and is processed into different allays of iron. What is the oxidation number of carbon in \( \text{CaCO}_3 \)? [1]

12. Base your answer to the question on the information below.
Element \( X \) is a solid metal that reacts with chlorine to form a water-soluble binary compound. The binary compound consists of element \( X \) and chlorine in a 1:2 molar ratio. What is the oxidation number of element \( X \) in this compound?

13. What is the oxidation number of nitrogen in \( \text{NO}(g) \)? [1]

14. Base your answer on the information below.
The catalytic converter in an automobile changes harmful gases produced during fuel combustion to less harmful exhaust gases. In the catalytic converter, nitrogen dioxide reacts with carbon monoxide to produce nitrogen and carbon dioxide. In addition, some carbon monoxide reacts with oxygen, producing carbon dioxide in the converter. These reactions are represented by the balanced equations below.

Reaction 1:
\[ 2\text{NO}_2(g) + 4\text{CO}(g) \rightarrow \text{N}_2(g) + 4\text{CO}_2(g) + 1198.4\text{ kJ} \]

Reaction 2:
\[ 2\text{CO}(g) + \text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 566.0\text{ kJ} \]

Determine the oxidation number of carbon in each carbon compound in reaction 2. Your response must include both the sign and value of each oxidation number. [1]

**Table J activity/ spontaneous reactions:**

1. Which metal can replace \( \text{Cr} \) in \( \text{Cr}_2\text{O}_3 \)?
   1. nickel  
   2. lead  
   3. copper  
   4. aluminum

2. Which metal is more active than \( \text{H}_2 \)?
   1. Ag  
   2. Au  
   3. Cu  
   4. Pb
3. Which metal reacts spontaneously with a solution containing zinc ions?
   1. magnesium
   2. nickel
   3. copper
   4. silver

4. Two chemistry students each combine a different metal with hydrochloric acid. Student A uses zinc, and hydrogen gas is readily produced. Student B uses copper, and no hydrogen gas is produced. Using Reference Table J, identify another metal that will react with hydrochloric acid to yield hydrogen gas. [1]

5. Fe(s) + 2HNO₃(aq) -> Fe(NO₃)₂(aq) + H₂(g)
Explain, using information from Reference Table J, why this reaction is spontaneous.

6. A flashlight can be powered by a rechargeable nickel-cadmium battery. In the battery, the anode is Cd(s) and the cathode is NiO₂(s). The unbalanced equation below represents the reaction that occurs as the battery produces electricity. When a nickel-cadmium battery is recharged, the reverse reaction occurs (see accompanying equation). Explain why Cd would be above Ni if placed on Table J. [1]

7. In a laboratory investigation, a student constructs a voltaic cell with iron and copper electrodes. Another student constructs a voltaic cell with zinc and iron electrodes. Testing the cells during operation enables the students to write the balanced ionic equations below.
   Cell with iron and copper electrodes:
   Cu²⁺(aq) + Fe(s) → Cu(s) + Fe²⁺(aq)
   Cell with zinc and iron electrodes:
   Fe²⁺(aq) + Zn(s) → Fe(s) + Zn²⁺(aq)
   State the relative activity of the three metals used in these two voltaic cells. [1]

8. Base your answer on the accompanying diagram of a voltaic cell and the balanced ionic equation shown. Identify one metal from Reference Table J that is more easily oxidized than Mg(s).

9. Which reaction occurs spontaneously?
   1. Cl₂(g) + 2NaBr(aq) → Br₂(l) + 2NaCl(aq)
   2. Cl₂(g) + 2NaF(aq) → F₂(g) + 2NaCl(aq)
   3. I₂(s) + 2NaBr(aq) --> Br₂(l) + 2NaI(aq)
   4. I₂(s) + 2NaF(aq) --> F₂(g) + 2NaI(aq)
Redox reactions and half reactions:

1. Which expression correctly represents a balanced reduction half-reaction?
   
   \[ (1) \ Na^+ + e^- \rightarrow Na \]
   
   \[ (2) \ Na \rightarrow Na^+ + e^- \]
   
   \[ (3) \ Cl_2 + 2e^- \rightarrow Cl^- \]
   
   \[ (4) \ 2Cl^- \rightarrow Cl_2 + 2e^- \]

2. Given the reaction: 
   \[ 2Na(s) + 2H_2O(\ell) \rightarrow 2NaOH(aq) + H_2(g) \]
   Which substance undergoes oxidation?
   
   1. Na
   2. NaOH
   3. H_2
   4. H_2O

3. As the reaction occurs, what happens to copper?
   
   \[ Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(\ell) \]
   
   1. It undergoes reduction and its oxidation number decreases.
   2. It undergoes reduction and its oxidation number increases.
   3. It undergoes oxidation and its oxidation number decreases.
   4. It undergoes oxidation and its oxidation number increases.

4. In any redox reaction, a reactant can undergo a decrease in oxidation number by
   
   1. losing electrons, only
   2. gaining electrons, only
   3. losing protons, only
   4. gaining protons, only

5. Which is a redox reaction?
   
   \[ (1) \ H^+ + Cl^- \rightarrow HCl \]
   
   \[ (2) \ NaOH + HCl \rightarrow NaCl + H_2O \]
   
   \[ (3) \ Fe + 2HCl \rightarrow FeCl_2 + H_2 \]
   
   \[ (4) \ MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O \]

6. Given the nickel-cadmium battery reaction: . . . (see image)
   
   During the discharge of the battery, Ni^{3+} ions are
   
   \[ 2NiOOH + Cd + 2H_2O \xrightarrow{\text{discharge}} 2Ni(OH)_2 + Cd(OH)_2 \]
   
   1. reduced, and cadmium metal is reduced
   2. reduced, and cadmium metal is oxidized
   3. oxidized, and cadmium metal is reduced
   4. oxidized, and cadmium metal is oxidized

7. Given the reaction shown:
   
   Which species undergoes oxidation?
\[ \text{Mg}(s) + 2 \text{H}^+(aq) + 2 \text{Cl}^-(aq) \rightarrow \text{Mg}^{2+}(aq) + 2 \text{Cl}^-(aq) + \text{H}_2(g) \]

1. Mg(s) 3. Cl(aq)
2. H^+(aq) 4. H_2(g)

8. Which particles are gained and lost during a redox reaction?
   1. electrons 3. neutrons
   2. protons 4. positrons

9. Given the equation:
   
   \[ 2 \text{Al} + 3 \text{Cu}^{2+} \rightarrow 2 \text{Al}^{3+} + 3 \text{Cu} \]

   The reduction half-reaction is
   
   \[(1) \quad \text{Al} \rightarrow \text{Al}^{3+} + 3e^- \quad (3) \quad \text{Al} + 3e^- \rightarrow \text{Al}^{3+} \]
   \[(2) \quad \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \quad (4) \quad \text{Cu}^{2+} \rightarrow \text{Cu} + 2e^- \]

10. Given the reaction for the corrosion of aluminum: \( 4 \text{Al} + 3 \text{O}_2 \rightarrow 2 \text{Al}_2\text{O}_3 \)
    Which half-reaction correctly represents the oxidation that occurs?
    1. \( \text{Al} + 3e^- \rightarrow \text{Al}^{3+} \)
    2. \( \text{Al} \rightarrow \text{Al}^{3+} + 3e^- \)
    3. \( \text{O}_2 + 4e^- \rightarrow 2 \text{O}^{2-} \)
    4. \( \text{O}_2 \rightarrow 2 \text{O}^{2-} + 4e^- \)

11. Which balanced equation represents a redox reaction?
    1. \( \text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3 \)
    2. \( \text{BaCl}_2 + \text{K}_2\text{CO}_3 \rightarrow \text{BaCO}_3 + 2\text{KCl} \)
    3. \( \text{CuO} + \text{CO} \rightarrow \text{Cu} + \text{CO}_2 \)
    4. \( \text{HCl} + \text{KOH} \rightarrow \text{KCl} + \text{H}_2\text{O} \)

12. Which changes occur when Pt\(^{2+}\) is reduced?
    1. The Pt\(^{2+}\) gains electrons and its oxidation number increases.
    2. The Pt\(^{2+}\) gains electrons and its oxidation number decreases.
    3. The Pt\(^{2+}\) loses electrons and its oxidation number increases.
    4. The Pt\(^{2+}\) loses electrons and its oxidation number decreases.

13. Which balanced equation represents an oxidation-reduction reaction?
    1. \( \text{BaCl}_2 + \text{Na}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + 2\text{NaCl} \)
    2. \( \text{C} + \text{H}_2\text{O} \rightarrow \text{CO} + \text{H}_2 \)
    3. \( \text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2 \)
    4. \( \text{Mg(OH)}_2 + 2\text{HNO}_3 \rightarrow \text{Mg(NO}_3)_2 = 2\text{H}_2\text{O} \)

14. Given the balanced ionic equation: \( 2\text{Al}^{3+}(aq) + 3\text{Mg}(s) \rightarrow 3\text{Mg}^{2+}(aq) + 2\text{Al}(s) \)
    In this reaction, electrons are transferred from
    1. Al to Mg\(^{2+}\)
    2. Al\(^{3+}\) to Mg
    3. Mg to Al\(^{3+}\)
    4. Mg\(^{2+}\) to Al
15. Which half-reaction equation represents the reduction of a potassium ion?

1. \( K^+ + e^- \rightarrow K \)
2. \( K + e^- \rightarrow K^+ \)
3. \( K^+ \rightarrow K + e^- \)
4. \( K \rightarrow K^+ + e^- \)

16. Which half-reaction equation represents the reduction of an iron(II) ion?

1. \( Fe^{2+} \rightarrow Fe^{3+} + e^- \)
2. \( Fe^{2+} + 2e^- \rightarrow Fe \)
3. \( Fe^{3+} + e^- \rightarrow Fe^{2+} \)
4. \( Fe \rightarrow Fe^{2+} + 2e^- \)

17. Given the balanced ionic equation representing the reaction in an operating voltaic cell:
   \[ \text{Zn(s)} + \text{Cu}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu(s)} \]

The flow of electrons through the external circuit in this cell is from the

1. Cu anode to the Zn cathode
2. Cu cathode to the Zn anode
3. Zn anode to the Cu cathode
4. Zn cathode to the Cu anode

18. Which balanced equation represents an oxidation-reduction reaction?

1. \( (1) \text{Ba(NO}_3\text{)}_2 + \text{Na}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + 2\text{NaNO}_3 \)
2. \( (2) \text{H}_3\text{PO}_4 + 3\text{KOH} \rightarrow \text{K}_3\text{PO}_4 + 3\text{H}_2\text{O} \)
3. \( (3) \text{Fe(s)} + \text{S(s)} \rightarrow \text{FeS(s)} \)
4. \( (4) \text{NH}_3(g) + \text{HCl(g)} \rightarrow \text{NH}_4\text{Cl(s)} \)

19. Given the reaction: \( \text{Zn(s)} + 2 \text{HCl(aq)} \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g) \)

Which statement correctly describes what occurs when this reaction takes place in a closed system?

1. Atoms of Zn(s) lose electrons and are oxidized.
2. Atoms of Zn(s) gain electrons and are reduced.
3. There is a net loss of mass.
4. There is a net gain of mass.

20. Which reaction is an example of an oxidation-reduction reaction?

1. \( \text{AgNO}_3 + \text{KI} \rightarrow \text{AgI} + \text{KNO}_3 \)
2. \( \text{Cu} + 2 \text{AgNO}_3 \rightarrow \text{Cu(NO}_3\text{)}_2 + 2 \text{Ag} \)
3. \( 2 \text{KOH} + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + 2 \text{H}_2\text{O} \)
4. \( \text{Ba(OH)}_2 + 2 \text{HCl} \rightarrow \text{BaCl}_2 + 2 \text{H}_2\text{O} \)

21. Half-reactions can be written to represent all

1. double-replacement reactions
2. neutralization reactions
3. fission and fusion reactions
4. oxidation and reduction reactions

22. Given the balanced equation representing a redox reaction: \( 2\text{Al} + 3\text{Cu}^{2+} \rightarrow 2\text{Al}^{3+} + 3\text{Cu} \)

Which statement is true about this reaction?

1. Each Al loses 2e\(^-\) and each Cu\(^{2+}\) gains 3e\(^-\).
2. Each Al loses 3e\(^-\) and each Cu\(^{2+}\) gains 2e\(^-\).
3. Each Al\(^{3+}\) gains 2e\(^-\) and each Cu loses 3e\(^-\).
4. Each Al\(^{3+}\) gains 3e\(^-\) and each Cu loses 2e\(^-\).

23. Which half-reaction shows conservation of charge?

1. \( \text{Cu} + e^- \rightarrow \text{Cu}^+ \)
2. \( \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \)
3. \( \text{Cu}^+ \rightarrow \text{Cu} + e^- \)
4. \( \text{Cu}^{2+} \rightarrow \text{Cu} + 2e^- \)
24. Given the balanced equation representing the reaction occurring in a voltaic cell:
\[ \text{Zn(s)} + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Pb(s)} \]. In the completed external circuit, the electrons flow from

1. Pb(s) to Zn(s) 
2. Pb^{2+}(aq) to Zn^{2+}(aq) 
3. Zn(s) to Pb(s) 
4. Zn^{2+}(aq) to Pb^{2+}(aq)

25. Which balanced equation represents a redox reaction?

1. \( \text{CuCO}_3(\text{s}) \rightarrow \text{CuO}(\text{s}) + \text{CO}_2(\text{g}) \) 
2. \( 2\text{KClO}_3(\text{s}) \rightarrow 2\text{KCl}(\text{s}) + 3\text{O}_2(\text{g}) \) 
3. \( \text{AgNO}_3(\text{aq}) + \text{KCl}(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{KNO}_3(\text{aq}) \) 
4. \( \text{H}_2\text{SO}_4(\text{aq}) + 2\text{KOH}(\text{aq}) \rightarrow \text{K}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\ell) \)

26. Given the unbalanced ionic equation: \( 3\text{Mg} + \text{Fe}^{3+} \rightarrow 3\text{Mg}^{2+} + \text{Fe} \)

When this equation is balanced, both Fe^{3+} and Fe have a coefficient of

1. 1, because a total of 6 electrons is transferred
2. 2, because a total of 6 electrons is transferred
3. 1, because a total of 3 electrons is transferred
4. 2, because a total of 3 electrons is transferred

27. Given the balanced equation: \( \text{Mg(s)} + \text{Ni}^{2+}(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{Ni(s)} \)

What is the total number of moles of electrons lost by Mg(s) when 2.0 moles of electrons are gained by Ni^{2+}(aq)?

1. 1.0 mol
2. 2.0 mol
3. 3.0 mol
4. 4.0 mol

28. Which half-reaction correctly represents reduction?

\[
\begin{align*}
(1) & \quad \text{Mn}^{4+} & \rightarrow & \text{Mn}^{3+} + \text{e}^- \\
(2) & \quad \text{Mn}^{4+} & \rightarrow & \text{Mn}^{7+} + 3\text{e}^- \\
(3) & \quad \text{Mn}^{4+} + \text{e}^- & \rightarrow & \text{Mn}^{3+} \\
(4) & \quad \text{Mn}^{4+} + 3\text{e}^- & \rightarrow & \text{Mn}^{7+}
\end{align*}
\]

29. Which equation represents an oxidation-reduction reaction?

1. \( \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \)
2. \( \text{H}_2\text{SO}_4 + \text{Ca(OH)}_2 \rightarrow \text{CaSO}_4 + 2\text{H}_2\text{O} \)
3. \( \text{MgCrO}_4 + \text{BaCl}_2 \rightarrow \text{MgCl}_2 + \text{BaCrO}_4 \)
4. \( \text{Zn(NO}_3)_2 + \text{Na}_2\text{CO}_3 \rightarrow 2\text{NaNO}_3 + \text{ZnCO}_3 \)

30. Which balanced equation represents a redox reaction?

\[
\begin{align*}
(1) & \quad \text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{NaNO}_3(\text{aq}) \\
(2) & \quad \text{H}_3\text{CO}_3(\text{aq}) \rightarrow \text{H}_2\text{O}(\ell) + \text{CO}_2(\text{g}) \\
(3) & \quad \text{NaOH}(\text{aq}) + \text{HCl}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\ell) \\
(4) & \quad \text{Mg(s)} + 2\text{HCl}(\text{aq}) \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2(\text{g})
\end{align*}
\]
31. Base your answer on the following redox reaction, which occurs spontaneously in an electrochemical cell. In the spaces provided, balance the equation using the smallest whole-number coefficients. [1]

\[
\text{Zn} + \text{Cr}^{3+} \rightarrow \text{Zn}^{2+} + \text{Cr}
\]

32. Base your answer on the following redox reaction, which occurs spontaneously in an electrochemical cell.
Zn + Cr\(^{3+}\) \rightarrow Zn\(^{2+}\) + Cr
a) Which half-reaction occurs at the cathode? [1]
b) State what happens to the number of protons in a Zn atom when it changes to Zn\(^{2+}\) as the redox reaction occurs. [1]

33. Aluminum is one of the most abundant metals in Earth's crust. The aluminum compound found in bauxite ore is Al\(_2\)O\(_3\). Over one hundred years ago, it was difficult and expensive to isolate aluminum from bauxite ore. In 1886, a brother and sister team, Charles and Julia Hall, found that molten (melted) cryolite, Na\(_3\)AlF\(_6\), would dissolve bauxite ore. Electrolysis of the resulting mixture caused the aluminum ions in the Al\(_2\)O\(_3\) to be reduced to molten aluminum metal. This less expensive process is known as the Hall process. Write the balanced half-reaction equation for the reduction of Al\(^{3+}\) to Al.

34. Given the reaction shown: \(4 \text{ Al}(s) + 3 \text{ O}_2(g) \rightarrow 2 \text{ Al}_2\text{O}_3(s)\)
   a) Write the balanced oxidation half-reaction for this oxidation-reduction reaction. [1]
   b) What is the oxidation number of oxygen in Al\(_2\)O\(_3\)? [1]

35. Given the reaction: Cl\(_2\) + 2 HBr \rightarrow Br\(_2\) + 2 HCl
   On a separate sheet of paper write a correctly balanced reduction half-reaction for this equation. [1]

36. Base your answer to the question on the accompanying diagram and balanced equation, which represent the electrolysis of molten NaCl. Write the balanced half-reaction for the reduction that occurs in this electrolytic cell.

37. Rust on an automobile door contains Fe\(_2\)O\(_3\)(s). The balanced equation representing one of the reactions between iron in the door of the automobile and oxygen in the atmosphere is given below. 
\(4\text{Fe}(s) + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s)\)
Identify the type of chemical reaction represented by this equation. [1]
38. The accompanying diagram represents an operating voltaic cell at 298 K and 1.0 atmosphere in a laboratory investigation. The reaction occurring in the cell is represented by the balanced ionic equation shown.
a) Determine the total number of moles of Ni$^{2+}$(aq) ions produced when 4.0 moles of Ag$^{+}$(aq) ions completely react in this cell. [1]
b) Write a balanced half-reaction equation for the reduction that occurs in this cell. [1]

![Voltaic Cell Diagram]

2Ag$^{+}$(aq) + Ni(s) → 2Ag(s) + Ni$^{2+}$(aq)

39. A student performed a laboratory activity to observe the reaction between aluminum foil and an aqueous copper(II) chloride solution. The reaction is represented by the accompanying balanced equation.
$2$Al(s) + $3$CuCl$_2$(aq) → $3$Cu(s) + $2$AlCl$_3$(aq) + energy
The procedures and corresponding observations for the activities are given in the accompanying table. State one observation that indicates Cu$^{2+}$ ions became Cu atoms. [1]

<table>
<thead>
<tr>
<th>Procedure</th>
<th>Observation</th>
</tr>
</thead>
<tbody>
<tr>
<td>In a beaker, completely dissolve 5.00 g of CuCl$_2$ in 80.0 mL of H$_2$O.</td>
<td>• The solution is blue green.</td>
</tr>
<tr>
<td>Cut 1.5 g of Al(s) foil into small pieces. Add all the foil to the mixture in the beaker. Stir the contents for 1 minute.</td>
<td>• The surface of Al(s) foil appears partially black. • The beaker feels warm to the touch.</td>
</tr>
<tr>
<td>Observe the beaker and contents after 10 minutes.</td>
<td>• The liquid in the beaker appears colorless. • A reddish-brown solid is seen at the bottom of the beaker. • Some pieces of Al(s) with a partially black coating remain in the beaker.</td>
</tr>
</tbody>
</table>

40. Base your answer on the unbalanced redox reaction below.
$\text{Cu}(s) + \text{AgNO}_3$(aq) $\rightarrow$ $\text{Cu(NO}_3)_2$(aq) + Ag(s)
Write the reduction half-reaction. [1]
41. In a laboratory investigation, a student constructs a voltaic cell with iron and copper electrodes. Another student constructs a voltaic cell with zinc and iron electrodes. Testing the cells during operation enables the students to write the balanced ionic equations below.

Cell with iron and copper electrodes:
\[ \text{Cu}^{2+} (aq) + \text{Fe}(s) \rightarrow \text{Cu}(s) + \text{Fe}^{2+} (aq) \]

Cell with zinc and iron electrodes:
\[ \text{Fe}^{2+} (aq) + \text{Zn}(s) \rightarrow \text{Fe}(s) + \text{Zn}^{2+} (aq) \]

a) Identify the particles transferred between Fe\(^{2+}\) and Zn during the reaction in the cell with zinc and iron electrodes. [1]

b) Write a balanced half-reaction equation for the reduction that takes place in the cell with zinc and iron electrodes. [1]

**Electrochemical cells:**

1. Base your answer on the information below.
   Electroplating is an electrolytic process used to coat metal objects with a more expensive and less reactive metal. The accompanying diagram shows an electroplating cell that includes a battery connected to a silver bar and a metal spoon. The bar and spoon are submerged in AgNO\(_3\)(aq). Explain why AgNO\(_3\) is a better choice than AgCl for use in this electrolytic process. [1]

[An Electroplating Cell diagram]

2. Which component of an electrochemical cell is correctly paired with its function?
   1. external conductor -- allows the solutions to mix
   2. external conductor -- permits the migration of ions
   3. salt bridge -- allows the solutions to mix
   4. salt bridge -- permits the migration of ions

3. Which procedure requires the use of an external electric current to force a redox reaction to occur?
   1. polymerization
   2. distillation
   3. electrolysis
   4. saponification
4. Which process occurs at the anode in an electrochemical cell?
   1. the loss of protons
   2. the loss of electrons
   3. the gain of protons
   4. the gain of electrons

5. Which energy conversion occurs during the operation of a voltaic cell?
   1. Chemical energy is spontaneously converted to electrical energy.
   2. Chemical energy is converted to electrical energy only when an external power source is provided.
   3. Electrical energy is spontaneously converted to chemical energy.
   4. Electrical energy is converted to chemical energy only when an external power source is provided.

6. Which statement describes one characteristic of an operating electrolytic cell?
   1. It produces electrical energy.
   2. It requires an external energy source.
   3. It uses radioactive nuclides.
   4. It undergoes a spontaneous redox reaction.

7. Which statement describes where the oxidation and reduction half-reactions occur in an operating electrochemical cell?
   1. Oxidation and reduction both occur at the anode.
   2. Oxidation and reduction both occur at the cathode.
   3. Oxidation occurs at the anode, and reduction occurs at the cathode.
   4. Oxidation occurs at the cathode, and reduction occurs at the anode.

8. Which energy conversion occurs in an operating electrolytic cell?
   1. chemical energy to electrical energy
   2. electrical energy to chemical energy
   3. nuclear energy to thermal energy
   4. thermal energy to nuclear energy

9. A voltaic cell differs from an electrolytic cell in that in a voltaic cell
   1. energy is produced when the reaction occurs
   2. energy is required for the reaction to occur
   3. both oxidation and reduction occur
   4. neither oxidation nor reduction occurs

10. What is the purpose of the salt bridge in a voltaic cell?
    1. It blocks the flow of electrons.
    2. It blocks the flow of positive and negative ions.
    3. It is a path for the flow of electrons.
    4. It is a path for the flow of positive and negative ions.

11. In a voltaic cell, chemical energy is converted to
    1. electrical energy, spontaneously
    2. electrical energy, nonspontaneously
    3. nuclear energy, spontaneously
    4. nuclear energy, nonspontaneously
12. Which statement is true about oxidation and reduction in an electrochemical cell?

1. Both occur at the anode.
2. Both occur at the cathode.
3. Oxidation occurs at the anode and reduction occurs at the cathode.
4. Oxidation occurs at the cathode and reduction occurs at the anode.

13. Which conversion of energy always occurs in a voltaic cell?

1. light energy to chemical energy
2. electrical energy to chemical energy
3. chemical energy to light energy
4. chemical energy to electrical energy

14. A student collects the materials and equipment below to construct a voltaic cell.
   - two 250-mL beakers
   - wire and a switch
   - one strip of magnesium
   - one strip of copper
   - 125 mL of 0.20 M Mg(NO_3)_2(aq)
   - 125 mL of 0.20 M Cu(NO_3)_2(aq)
   Which additional item is required for the construction of the voltaic cell?

1. an anode
2. a battery
3. a cathode
4. a salt bridge

15. A voltaic cell spontaneously converts chemical energy to

1. electrical energy
2. geothermal energy
3. mechanical energy
4. nuclear energy

16. Given the balanced equation representing a reaction occurring in an electrolytic cell: (see image)
Where is Na(l) produced in the cell?

\[2\text{NaCl}(l) \rightarrow 2\text{Na}(l) + \text{Cl}_2(g)\]

1. at the anode, where oxidation occurs
2. at the anode, where reduction occurs
3. at the cathode, where oxidation occurs
4. at the cathode, where reduction occurs

17. Which energy conversion occurs during the operation of an electrolytic cell?

1. chemical energy to electrical energy
2. electrical energy to chemical energy
3. nuclear energy to electrical energy
4. electrical energy to nuclear energy

18. State one difference between voltaic cells and electrolytic cells. Include information about both types of cells in your answer. [1]

19. Base your answer on the accompanying diagram of a voltaic cell and the balanced ionic equation shown.
What is the total number of moles of electrons needed to completely reduce 6.0 moles of Ni^{2+}(aq) ions?
20. Base your answer on the accompanying diagram of a voltaic cell and the balanced ionic equation shown. Explain the function of the salt bridge in the voltaic cell.

21. Base your answer on the information below. Aluminum is one of the most abundant metals in Earth's crust. The aluminum compound found in bauxite ore is Al₂O₃. Over one hundred years ago, it was difficult and expensive to isolate aluminum from bauxite ore. In 1886, a brother and sister team, Charles and Julia Hall, found that molten (melted) cryolite, Na₃AlF₆, would dissolve bauxite ore. Electrolysis of the resulting mixture caused the aluminum ions in the Al₂O₃ to be reduced to molten aluminum metal. This less expensive process is known as the Hall process.

22. Base your answer on the information below. A flashlight can be powered by a rechargeable nickel-cadmium battery. In the battery, the anode is Cd(s) and the cathode is NiO₂(s). The unbalanced equation below represents the reaction that occurs as the battery produces electricity. When a nickel-cadmium battery is recharged, the reverse reaction occurs (see accompanying equation).

23. Base your answer on the information below. A voltaic cell with magnesium and copper electrodes is shown in the accompanying diagram. The copper electrode has a mass of 15.0 grams. When the switch is closed, the reaction in the cell begins. The balanced ionic equation for the reaction in the cell is shown below the cell diagram. After several hours, the copper electrode is removed, rinsed with water, and dried. At this time, the mass of the copper electrode is greater than 15.0 grams.

   a) State the direction of electron flow through the wire between the electrodes when the switch is closed. [1]
   b) State the purpose of the salt bridge in this cell. [1]
   c) Explain, the terms of copper ions and copper atoms, why the mass of the copper electrode increases as the cell operates. Your response must include information about both copper ions and copper atoms. [1]
24. Base your answer to the question on the accompanying diagram and balanced equation, which represent the electrolysis of molten NaCl.

a) When the switch is closed, which electrode will attract the sodium ions?

b) What is the purpose of the battery in this electrolytic cell?

25. Base your answer on the information below.
In a laboratory investigation, a student constructs a voltaic cell with iron and copper electrodes. Another student constructs a voltaic cell with zinc and iron electrodes. Testing the cells during operation enables the students to write the balanced ionic equations below.
Cell with iron and copper electrodes:
Cu^{2+}(aq) + Fe(s) → Cu(s) + Fe^{2+}(aq)

Cell with zinc and iron electrodes:
Fe^{2+}(aq) + Zn(s) → Fe(s) + Zn^{2+}(aq)

State evidence from the balanced equation for the cell with iron and copper electrodes that indicates the reaction in the cell is an oxidation-reduction reaction. [1]

26. Base your answer on the information below.
Electroplating is an electrolytic process used to coat metal objects with a more expensive and less reactive metal. The accompanying diagram shows an electroplating cell that includes a battery connected to a silver bar and a metal spoon. The bar and spoon are submerged in AgNO₃(aq).

Explain the purpose of the battery in this cell. [1]
27. Base your answer on the accompanying information.
The accompanying diagram represents an operating voltaic cell at 298 K and 1.0 atmosphere in a
laboratory investigation. The reaction occurring in the cell is represented by the balanced ionic
equation shown.
Identify the anode in this cell. [1]

Organic in general:

1. Which element must be present in an organic compound?
   1. hydrogen  2. oxygen  3. carbon  4. nitrogen

2. Which element has atoms that can bond with each other to form long chains or rings?
   1. carbon  2. nitrogen  3. oxygen  4. fluorine
3. Given the formulas for two compounds, these compounds differ in:

\[
\begin{align*}
\text{H}_3\text{C}-\text{C}=\text{O}-\text{C}-\text{C}-\text{H} \\
\text{H}&\quad\text{H} & \text{H} & \text{H} & \text{H} \\
\text{H}&\quad\text{H} & \text{H} & \text{H} & \text{H} \\
\text{H}-\text{C}-\text{C}-\text{C}-\text{OH} \\
\text{H}&\quad\text{H} & \text{H} & \text{H} & \text{H}
\end{align*}
\]

1. gram-formula mass  
2. molecular formula  
3. percent composition by mass  
4. physical properties at STP

4. Atoms of which element can bond to each other to form chains, rings, and networks?

1. carbon  
2. fluorine  
3. hydrogen  
4. oxygen

5. A molecule of an organic compound contains at least one atom of:

1. carbon  
2. chlorine  
3. nitrogen  
4. oxygen

6. Which atoms can bond with each other to form chains, rings, or networks?

1. carbon atoms  
2. hydrogen atoms  
3. oxygen atoms  
4. nitrogen atoms

7. Which structural formula is incorrect?

\[
\begin{align*}
\text{H}&\quad\text{H} & \text{C}=\text{C} & \text{H} \\
\text{H}&\quad\text{H} & \text{H} & \text{H}
\end{align*}
\]

(1)  

\[
\begin{align*}
\text{O} \\
\text{H}&\quad\text{C}=\text{C} & \text{H} \\
\text{H}&\quad\text{H} & \text{H} & \text{H} & \text{H}
\end{align*}
\]

(2)  

(3)  

(4)

8. Atoms of which element can bond with each other to form ring and chain structures in compounds?

1. C  
2. Ca  
3. H  
4. Na

9. Which element is present in all organic compounds?

1. carbon  
2. hydrogen  
3. nitrogen  
4. oxygen
10. Base your answer on the accompanying equation, which represents an organic compound reacting with bromine.
What is the gram-formula mass of the product in this reaction? [1]

\[
\text{H}_3\text{C}=\text{C}-\text{H} + \text{Br}_2 \rightarrow \text{H}-\text{C}-\text{C}-\text{H} \quad \text{Br} \quad \text{Br} \quad \text{H}
\]

11. Base your answer on the information below and accompanying diagram and on your knowledge of chemistry.
Crude oil is a mixture of many hydrocarbons that have different numbers of carbon atoms. The use of a fractionating tower allows the separation of this mixture based on the boiling points of the hydrocarbons.
To begin the separation process, the crude oil is heated to about 400°C in a furnace, causing many of the hydrocarbons of the crude oil to vaporize. The vaporized mixture is pumped into a fractionating tower that is usually more than 30 meters tall. The temperature of the tower is highest at the bottom. As vaporized samples of hydrocarbons travel up the tower, they cool and condense. The liquid hydrocarbons are collected on trays and removed from the tower. The accompanying diagram illustrates the fractional distillation of the crude oil and the temperature ranges in which the different hydrocarbons condense.
State the trend between the boiling point of the hydrocarbons contained in the crude oil and the number of carbon atoms in these molecules.

**Hydrocarbons and homologous series / Table Q:**

1. Which compound is a saturated hydrocarbon?
   1. hexane
   2. hexene
   3. hexanol
   4. hexanal
2. A double carbon-carbon bond is found in a molecule of
   1. pentane
   2. pentene
   3. pentyne
   4. pentanol

3. Which compound is an unsaturated hydrocarbon?
   1. hexanal
   2. hexane
   3. hexanoic acid
   4. hexyne

4. Which general formula represents the compound CH₃CH₂CCH?
   1. CₙHₙ
   2. CₙH₂ₙ
   3. CₙH₂ₙ - 2
   4. CₙH₂ₙ + 2

5. Which compound is a saturated hydrocarbon?
   1. propanal
   2. propane
   3. propene
   4. propyne

6. Which formula represents 2-butene?

\[ \begin{align*}
   &\text{(1)} & \text{(2)} & \text{(3)} & \text{(4)} \\
   &\text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\
   &\text{H} & \text{C} & - & \text{C} & - & \text{C} & - & \text{H} \\
   &\text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\
   &\text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\
\end{align*} \]

7. A carbon-carbon triple bond is found in a molecule of
   1. butane
   2. butanone
   3. butene
   4. butyne

8. A molecule of an unsaturated hydrocarbon must have
   1. at least one single carbon-carbon bond
   2. at least one multiple carbon-carbon bond
   3. two or more single carbon-carbon bonds
   4. two or more multiple carbon-carbon bonds

9. Which organic compound is a saturated hydrocarbon?
   1. ethyne
   2. ethene
   3. ethanol
   4. ethane
10. Which formula represents an unsaturated hydrocarbon?

- (1) \[ \text{H} - \text{C} - \text{C} - \text{H} \]
- (3) \[ \text{H} - \text{C} - \text{C} = \text{C} - \text{H} \]
- (2) \[ \text{H} - \text{C} - \text{C} - \text{C} - \text{H} \]
- (4) \[ \text{H} - \text{C} - \text{O} - \text{C} - \text{C} - \text{H} \]

11. Which structural formula correctly represents a hydrocarbon molecule?

- (1) \[ \text{H} - \text{C} - \text{C} - \text{H} \]
- (3) \[ \text{H} - \text{C} - \text{C} = \text{O} \]
- (2) \[ \text{H} - \text{C} = \text{C} - \text{H} \]
- (4) \[ \text{H} - \text{C} = \text{C} - \text{H} \]

12. What is the IUPAC name of the compound with the structural formula shown?

- 1. 2-pentene
- 2. 3-pentene
- 3. 2-pentyne
- 4. 3-pentyne

13. A molecule of butane and a molecule of 2-butene both have the same total number of

- 1. carbon atoms
- 2. hydrogen atoms
- 3. single bonds
- 4. double bonds

14. Which general formula represents the homologous series of hydrocarbons that includes the compound 1-heptyne?

- 1. \( \text{C}_n \text{H}_{2n-6} \)
- 2. \( \text{C}_n \text{H}_{2n-2} \)
- 3. \( \text{C}_n \text{H}_{2n} \)
- 4. \( \text{C}_n \text{H}_{2n+2} \)

15. Which formula represents an unsaturated hydrocarbon?
16. Hydrocarbons are compounds that contain

1. carbon, only
2. carbon and hydrogen, only
3. carbon, hydrogen, and oxygen, only
4. carbon, hydrogen, oxygen, and nitrogen, only

17. What is the IUPAC name of the organic compound that has the formula shown below?

1. 1,1-dimethylbutane
2. 2-methylpentane
3. hexane
4. 4-methylpentane

18. Which compound is a member of the same homologous series as C₃H₈?

1. CH₄
2. C₄H₈
3. C₅H₁₀
4. C₆H₁₄

19. A straight-chain hydrocarbon that has only one double bond in each molecule has the general formula

1. CₖH₂ₙ₋₆
2. CₖH₂ₙ₋₂
3. CₖH₂ₙ
4. CₖH₂ₙ₊₂

20. Which formula represents an unsaturated hydrocarbon?

1. C₃H₁₂
2. C₆H₁₄
3. C₇H₁₆
4. C₈H₁₄

21. Which formula represents a molecule of a saturated hydrocarbon?

1. C₂H₂
2. C₄H₁₀
3. C₅H₈
4. C₆H₆
22. Which formula represents an unsaturated hydrocarbon?

\[ \text{(1)} \quad \text{H} - \text{C} = \text{H} \quad \text{H} - \text{C} = \text{H} - \text{H} \\
\text{(2)} \quad \text{H} - \text{C} = \text{H} - \text{H} - \text{H} \quad \text{H} - \text{C} = \text{OH} - \text{H} \]

23. Base your answer on the information below and accompanying diagram and on your knowledge of chemistry.

Crude oil is a mixture of many hydrocarbons that have different numbers of carbon atoms. The use of a fractionating tower allows the separation of this mixture based on the boiling points of the hydrocarbons.

To begin the separation process, the crude oil is heated to about 400°C in a furnace, causing many of the hydrocarbons of the crude oil to vaporize. The vaporized mixture is pumped into a fractionating tower that is usually more than 30 meters tall. The temperature of the tower is highest at the bottom. As vaporized samples of hydrocarbons travel up the tower, they cool and condense. The liquid hydrocarbons are collected on trays and removed from the tower. The accompanying diagram illustrates the fractional distillation of the crude oil and the temperature ranges in which the different hydrocarbons condense.

a) Write an IUPAC name of one saturated hydrocarbon that leaves the fractionating tower at less than 40°C.

b) How many hydrogen atoms are present in one molecule of octane?
24. To which homologous series does \( \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3 \) belong?

25. Base your answer on the information below.
A reaction between bromine and a hydrocarbon is represented by the balanced accompanying equation.
Write the name of the homologous series to which the hydrocarbon belongs. [1]

\[
\text{Br}_2 + \text{H} - \text{C} = \text{C} - \text{C} - \text{H} \rightarrow \text{H} - \text{C} = \text{C} - \text{C} - \text{H}
\]

26. Base your answer on the accompanying equation, which represents an organic compound reacting with bromine.
What is the IUPAC name for the organic compound that reacts with \( \text{Br}_2 \)? [1]

\[
\text{H} - \text{C} = \text{C} - \text{C} - \text{H} + \text{Br}_2 \rightarrow \text{H} - \text{C} = \text{C} - \text{C} - \text{H}
\]

27. Base your answer on the information below.
Gasoline is a mixture composed primarily of hydrocarbons such as isooctane, which is also known as 2,2,4-trimethylpentane.
Gasoline is assigned a number called an octane rating. Gasoline with an octane rating of 87 performs the same as a mixture that consists of 87% isooctane and 13% heptane.
An alternative fuel, E-85, can be used in some automobiles. This fuel is a mixture of 85% ethanol and 15% gasoline.
In the space provided or on a separate piece of paper, draw a structural formula for a molecule of 2,2,4-trimethylpentane. [1]

Functional groups and compounds / Table R:

1. The organic compound represented by the condensed structural formula \( \text{CH}_3\text{CH}_2\text{CH}_2\text{CHO} \) is classified as an
   1. alcohol
   2. aldehyde
   3. ester
   4. ether

2. What is the total number of carbon atoms in a molecule of ethanoic acid?
   1. 1
   2. 2
   3. 3
   4. 4

3. Base your answer on the information below. The equation shown (see image) represents the reaction between butanoic acid and an unidentified reactant, \( X \).
Write the molecular formula of the organic product in the equation. [1]

\[
\text{H} - \text{C} = \text{C} - \text{C} - \text{C} - \text{OH} + X \rightarrow \text{H} - \text{C} - \text{C} - \text{C} - \text{O} - \text{C} - \text{H} + \text{H} - \text{O} - \text{H}
\]
4. Which Lewis electron-dot diagram represents chloroethene?

\[
\begin{align*}
(1) & \quad \text{H} : \text{C} : \text{C} : \text{Cl} : \\
(2) & \quad \text{H} : \text{C} : \text{C} : \text{Cl} : \\
(3) & \quad \text{H} \quad \text{H} \\
(4) & \quad \text{H} \quad \text{H}
\end{align*}
\]

5. Base your answer on the article below and on your knowledge of chemistry.

**The Decaffeinating Tradition**

For coffee beans to be labeled "decaffeinated," at least 97% of the caffeine must be removed. There are three primary methods for decaffeination: chemical extraction, the Swiss water process, and supercritical fluid extraction.

Although all methods of decaffeinating coffee involve the use of "chemicals," one process has been traditionally referred to as "chemical extraction," probably because it uses organic solvents that are not typically part of our normal environment. The traditional method offers two slightly varied options using dichloromethane (CH$_2$Cl$_2$) or ethyl acetate (CH$_3$COOC$_2$H$_5$) as solvents. With both solvents, the beans are first soaked in water to soften them and speed the decaffeinating process. The beans are then soaked in one of the two solvents, which dissolves the caffeine in the bean. Once the solvent has removed the caffeine, the coffee beans are treated with steam. This evaporates the organic solvent along with the caffeine.

The process is identical for both solvents, but many coffee companies prefer to use ethyl acetate to decaffeinate their coffee. This allows them to label the beans "naturally decaffeinated," because ethyl acetate occurs naturally in orange rinds and many other fruits. Although consumers may prefer this label, it is misleading. The ethyl acetate used is actually synthesized; it is not extracted from fruits because it would be too costly.

Both of these commercial methods have a growing number of detractors, prompting many coffee companies to turn toward other methods. Opposition arises because the solvents used can never be completely removed from the coffee beans. The traces left behind, however, are below the amounts required for the "decaffeinated" label. Because of the recognized potential hazards associated with the use of dichloromethane and ethyl acetate, the United States Food and Drug Administration and the United States Department of Agriculture continue to investigate and evaluate any possible dangers that might be associated with the use of these chemicals.

a. Draw the structural formula for ethyl acetate. (The correct IUPAC name is ethyl ethanoate.) [1]

b. To what class of organic compounds does ethyl acetate belong? [1]

c. Draw a correct structural formula for dichloromethane. [1]

d. To what class of organic compounds does dichloromethane belong? [1]

e. Which do you think is better, ethyl acetate or dichloromethane, as a decaffeinating agent for coffee? Explain your choice in terms of the information found in the article. [1]

f. Indicate whether you think the caffeine molecule is polar or nonpolar; then explain your answer in terms of the solubility of the caffeine molecule. [1]
6. Which class of organic compounds has molecules that contain nitrogen atoms?
   1. alcohol
   2. amine
   3. ether
   4. ketone

7. Given a formula of a functional group:

   \[ \text{O} \]
   \[ \text{C} \text{--OH} \]

   An organic compound that has this functional group is classified as

   1. an acid
   2. an aldehyde
   3. an ester
   4. a ketone

8. Given the structural formulas for two organic compounds (see image):
   The differences in their physical and chemical properties are primarily due to their different
   1. number of carbon atoms
   2. number of hydrogen atoms
   3. molecular masses
   4. functional groups

9. Which structural formula represents an alcohol?

10. Molecules of 1-bromopropane and 2-bromopropane differ in
    1. molecular formula
    2. structural formula
    3. number of carbon atoms per molecule
    4. number of bromine atoms per molecule

11. Given the three organic structural formulas shown (see image).
    Which organic-compound classes are represented by these structural formulas, as shown from left to right?
12. The compounds 2-butanol and 2-butene both contain
   1. double bonds, only
   2. single bonds, only
   3. carbon atoms
   4. oxygen atoms

13. Ethanol and dimethyl ether have different chemical and physical properties because they have different
   1. functional groups
   2. molecular masses
   3. numbers of covalent bonds
   4. percent compositions by mass

14. The molecule shown belongs to which class of compounds?

15. Base your answer on the information below. Given the reaction between 1-butene and chlorine gas:
   \[ C_4H_8 + Cl_2 \rightarrow C_4H_8Cl_2 \]
   Draw the structural formula of the product 1,2-dichlorobutane.

16. Base your answer on the information below.
   A thiol is very similar to an alcohol, but a thiol has a sulfur atom instead of an oxygen atom in the functional group. One of the compounds in a skunk's spray is 2-butene-1-thiol. The formula of this compound is shown in the accompanying diagram.
   Explain, in terms of composition, why this compound is a thiol.

17. Base your answer on the information below.
   An unlit candle is secured to the bottom of a 200-milliliter glass beaker. Baking soda (sodium hydrogen carbonate) is added around the base of the candle as shown below. The candle is lit and dilute ethanoic acid is poured down the inside of the beaker. As the acid reacts with the baking soda, bubbles of CO\(_2\) gas form. After a few seconds, the air in the beaker is replaced by 0.20 liter of
CO₂ gas, causing the candle flame to go out. The density of CO₂ gas is 1.8 grams per liter at room temperature.

In the space on the answer sheet or on a separate piece of paper, draw a structural formula for the acid that was poured into the beaker. [1]

18. Base your answer on the information below.
Glycine, NH₂CH₂COOH, is an organic compound found in proteins. Acetamide, CH₃CONH₂, is an organic compound that is an excellent solvent. Both glycine and acetamide consist of the same four elements, but the compounds have different functional groups.
Identify one functional group in a glycine molecule. [1]

Glycine, NH₂CH₂COOH, is an organic compound found in proteins. Acetamide, CH₃CONH₂, is an organic compound that is an excellent solvent. Both glycine and acetamide consist of the same four elements, but the compounds have different functional groups.
Draw a structural formula for acetamide. [1]

20. Given the ester: ethyl butanoate
   a In the space provided (on a separate sheet of paper), draw the structural formula for this ester. [1]  
   b Determine the gram formula mass of this ester. [1]

21. On a separate piece of paper draw the structural formula for butanoic acid. [1]

22. Base your answers to the question on the properties of propanone.
In the space provided or on a separate piece of paper, draw the structural formula for propanone.

23. Base your answer on the information below.
The incomplete equation shown represents an esterification reaction. The alcohol reactant is represented by X.
On the structural formula on the answer sheet, or, if taken online, on a separate piece of paper, circle the acid functional group, only. [1]

24. Base your answer on the information below.
The incomplete equation shown represents an esterification reaction. The alcohol reactant is represented by X.
a) Write an IUPAC name for the reactant represented by its structural formula in this equation. [1]
b) In the space on the answer sheet, or, if taken online, on a separate piece of paper, draw the structural formula for the alcohol represented by X. [1]

25. Base your answer on the information below. The equation shown (see image) represents the reaction between butanoic acid and an unidentified reactant, X.
In the space on the answer sheet or on a separate piece of paper, draw a structural formula for the unidentified reactant, X, in the equation. [1]

26. Base your answer on the information below.
Biodiesel is an alternative fuel for vehicles that use petroleum diesel. Biodiesel is produced by reacting vegetable oil with CH₃OH. Methyl palmitate, C₁₅H₃₁COOCH₃, a compound found in biodiesel, is made from soybean oil. One reaction of methyl palmitate with oxygen is represented by the balanced equation below.
2C₁₅H₃₁COOCH₃ + 49O₂ → 34CO₂ + 34H₂O + energy
a) Write an IUPAC name for the compound that reacts with vegetable oil to produce biodiesel. [1]
b) Identify the class of organic compounds to which methyl palmitate belongs. [1]

27. Base your answer on the information below.
Gasoline is a mixture composed primarily of hydrocarbons such as isooctane, which is also known as 2,2,4-trimethylpentane.
Gasoline is assigned a number called an octane rating. Gasoline with an octane rating of 87 performs the same as a mixture that consists of 87% isooctane and 13% heptane.
An alternative fuel, E-85, can be used in some automobiles. This fuel is a mixture of 85% ethanol and 15% gasoline.
Identify the functional group in a molecule of ethanol in the alternative fuel E-85. [1]

Isomers:

1. Which compound is an isomer of pentane?
   1. butane 3. methyl butane
   2. propane 4. methyl propane

2. Which compounds are isomers?
   1. CH₃OH and CH₂CH₂OH 3. CH₃CH₂CHO and CH₃COCH₃
   2. CH₄ and CCl₄ 4. CH₃CH₂OH and CH₃CH₂COOH
3. Which formula is a \( n \) isomer of butane?

\[
\begin{align*}
(1) & \quad \text{H} & \text{C} & \text{C} & \text{C} & \text{H} \\
(2) & \quad \text{H} & \text{C} & \text{C} & \text{C} & \text{H} \\
(3) & \quad \text{H} & \text{C} & \text{C} & \text{C} & \text{H} \\
(4) & \quad \text{H} & \text{C} & \text{C} & \text{C} & \text{H}
\end{align*}
\]

4. Which two compounds have the same molecular formula but different chemical and physical properties?

1. \( \text{CH}_3\text{CH}_2\text{Cl} \) and \( \text{CH}_3\text{CH}_2\text{Br} \)
2. \( \text{CH}_2\text{CHCH}_2 \) and \( \text{CH}_3\text{CH}_2\text{CH}_3 \)
3. \( \text{CH}_3\text{CHO} \) and \( \text{CH}_3\text{COCH}_3 \)
4. \( \text{CH}_3\text{CH}_2\text{OH} \) and \( \text{CH}_3\text{OCH}_3 \)

5. The isomers butane and methylpropane differ in their

1. molecular formulas
2. structural formulas
3. total number of atoms per molecule
4. total number of bonds per molecule

6. Given a formula representing a compound:
   Which formula represents an isomer of this compound?

\[
\begin{align*}
(1) & \quad \text{H} & \text{C} & \text{C} & \text{C} & \text{H} \\
(2) & \quad \text{H} & \text{C} & \text{C} & \text{C} & \text{H} \\
(3) & \quad \text{H} & \text{C} & \text{C} & \text{C} & \text{H} \\
(4) & \quad \text{H} & \text{C} & \text{C} & \text{C} & \text{H}
\end{align*}
\]
7. Which structural formula represents a molecule that is not an isomer of pentane?

(1) 

(2) 

(3) 

(4) 

8. Given the structural formula shown:

(a) \( \text{HCCCH} \) 
(b) \( \text{HCCCOH} \) 
(c) \( \text{HCCCH} \) 
(d) \( \text{HCCCOH} \) 

Which two formulas represent compounds that are isomers of each other?

1. \( A \) and \( B \) 
2. \( A \) and \( C \) 
3. \( B \) and \( D \) 
4. \( C \) and \( D \) 

9. Which two compounds are isomers of each other?

1. \( \text{CH}_3\text{CH}_2\text{COOH} \) and \( \text{CH}_3\text{COOCH}_2\text{CH}_3 \)
2. \( \text{CH}_3\text{CH}_2\text{CHO} \) and \( \text{CH}_3\text{COCH}_3 \)
3. \( \text{CH}_3\text{CHBrCH}_3 \) and \( \text{CH}_2\text{BrCHBrCH}_3 \)
4. \( \text{CH}_3\text{CHOHCH}_3 \) and \( \text{CH}_3\text{CHOHCH}_2\text{OH} \)

10. Molecules of 2-methyl butane and 2,2-dimethyl propane have different

1. structural formulas 
2. molecular formulas 
3. numbers of carbon atoms 
4. numbers of covalent bonds 

11. Given the structural formula for butane:

On a separate sheet of paper, draw the structural formula of an isomer of butane. [1]
12. Base your answer on the information below.
Biodiesel is an alternative fuel for vehicles that use petroleum diesel. Biodiesel is produced by
reacting vegetable oil with CH₃OH. Methyl palmitate, C₁₅H₃₁COOCH₃, a compound found in
biodiesel, is made from soybean oil. One reaction of methyl palmitate with oxygen is represented by
the balanced equation below.
2C₁₅H₃₁COOCH₃ + 49O₂ → 34CO₂ + 34H₂O + energy
Explain, in terms of both atoms and molecular structure, why there is no isomer of CH₃OH. [1]

Organic reactions:

1. When butane burns in an excess of oxygen, the principal products are
   1. CO₂ and H₂O
   2. CO₂ and H₂
   3. CO and H₂O
   4. CO and H₂

2. The process of joining many small molecules into larger molecules is called
   1. neutralization
   2. polymerization
   3. saponification
   4. substitution

3. Which type of reaction is represented by the equation (see image)?
   **Note:** n and m are very large numbers equal to
   about 2000.

   ![Reaction Diagram]

   1. esterification
   2. fermentation
   3. saponification
   4. polymerization

4. This reaction is an example of

   \[
   \text{CH₃C-OH + HOCl₂} \rightleftharpoons \text{CH₃C-O-C₂H₅ + H₂O}
   \]

   1. fermentation
   2. saponification
   3. hydrogenation
   4. esterification

5. Given the equation: \(X + \text{Cl}_2 \rightarrow \text{C}_2\text{H}_5\text{Cl} + \text{HCl}\). Which molecule is represented by \(X\)?
   1. C₂H₄
   2. C₂H₆
   3. C₃H₆
   4. C₃H₈

6. Given the incomplete equation representing an organic addition reaction: \(X(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow X\text{Cl}_2(\text{g})\)
   Which compound could be represented by \(X\)?
   1. CH₄
   2. C₂H₄
   3. C₃H₈
   4. C₄H₁₀
7. The reaction that joins thousands of small, identical molecules to form one very long molecule is called
   1. esterification
   2. fermentation
   3. polymerization
   4. substitution

8. Which reaction results in the production of soap?
   1. esterification
   2. fermentation
   3. polymerization
   4. saponification

9. The reaction between an organic acid and an alcohol produces
   1. an aldehyde
   2. a ketone
   3. an ether
   4. an ester

10. In the space to the right of the reactants and arrow provided (print this page), draw the structural formula for the product of the reaction shown. [1]

11. Base your answer on the information below. Given the reaction between 1-butene and chlorine gas: \( \text{C}_4\text{H}_8 + \text{Cl}_2 \rightarrow \text{C}_4\text{H}_8\text{Cl}_2 \)
    Which type of chemical reaction is represented by this equation?

12. Ozone gas, \( \text{O}_3 \), can be used to kill adult insects in storage bins for grain without damaging the grain. The ozone is produced from oxygen gas, \( \text{O}_2 \), in portable ozone generators located near the storage bins. The concentrations of ozone used are so low that they do not cause any environmental damage. This use of ozone is safer and more environmentally friendly than a method that used bromomethane, \( \text{CH}_3\text{Br} \). However, bromomethane was more effective than ozone because \( \text{CH}_3\text{Br} \) killed immature insects as well as adult insects.

   Adapted From: The Sunday Gazette (Schenectady, NY) 3/9/03

   Given the balanced equation for producing bromomethane:
   \( \text{Br}_2 + \text{CH}_4 \rightarrow \text{CH}_3\text{Br} + \text{HBr} \)
   Identify the type of organic reaction shown. [1]

13. During a bread-making process, glucose is converted to ethanol and carbon dioxide, causing the bread dough to rise. Zymase, an enzyme produced by yeast, is a catalyst needed for this reaction. Balance the accompanying equation in the space provided or on a separate piece of paper for the reaction that causes bread dough to rise, using the smallest whole-number coefficients. [1]
   b. In the space provided or on a separate piece of paper, draw a structural formula for the alcohol formed in this reaction. [1]
   c. State the effect of zymase on the activation energy for this reaction. [1]
14. Base your answer on the information below.
A reaction between bromine and a hydrocarbon is represented by the balanced accompanying equation. Identify the type of organic reaction. [1]

\[
\text{Br}_2 + H\text{-C} = \text{C} = \text{C} - \text{H} \rightarrow H\text{-C} = \text{C} = \text{C} - \text{H}
\]

15. Base your answer on the accompanying equation, which represents an organic compound reacting with bromine.
What type of organic reaction is represented by this equation? [1]

\[
\text{C} = \text{C} = \text{C} - \text{H} + \text{Br}_2 \rightarrow H\text{-C} = \text{C} = \text{C} - \text{H}
\]

16. Base your answer on the information below. The equation shown (see image) represents the reaction between butanoic acid and an unidentified reactant, X.
Identify the type of organic reaction represented by the equation. [1]

\[
\text{H} - \text{C} - \text{C} - \text{C} - \text{OH} + X \rightarrow \text{H} - \text{C} - \text{C} - \text{C} - \text{O} - \text{C} - \text{H} + \text{H} - \text{O} - \text{H}
\]

17. Base your answer on the information below.
Biodiesel is an alternative fuel for vehicles that use petroleum diesel. Biodiesel is produced by reacting vegetable oil with CH\(_3\)OH. Methyl palmitate, C\(_{15}\)H\(_{31}\)COOCH\(_3\), a compound found in biodiesel, is made from soybean oil. One reaction of methyl palmitate with oxygen is represented by the balanced equation below.

\[
2\text{C}_{15}\text{H}_{31}\text{COOCH}_3 + 49\text{O}_2 \rightarrow 34\text{CO}_2 + 34\text{H}_2\text{O} + \text{energy}
\]
Identify the type of organic reaction represented by the balanced equation. [1]

**Nuclear terms and particles:**

1. The change that is undergone by an atom of an element made radioactive by bombardment with high-energy protons is called
   1. natural transmutation
   2. artificial transmutation
   3. natural decay
   4. radioactive decay

2. Which type of reaction converts one element to another element?
   1. neutralization
   2. polymerization
   3. substitution
   4. transmutation

3. Which list of radioisotopes contains an alpha emitter, a beta emitter, and a positron emitter?
   1. C-14, N-16, P-32
   2. Cs-137, Fr-220, Tc-99
   3. Kr-85, Ne-19, Rn-222
   4. Pu-239, Th-232, U-238
4. Which nuclear decay emission consists of energy, only?
   1. alpha particle  
   2. beta particle  
   3. gamma radiation  
   4. positron

5. The energy released by a nuclear reaction results primarily from the
   1. breaking of bonds between atoms  
   2. formation of bonds between atoms  
   3. conversion of mass into energy  
   4. conversion of energy into mass

6. Which radioisotope is used in medicine to treat thyroid disorders?
   1. cobalt-60  
   2. iodine-131  
   3. phosphorus-32  
   4. uranium-238

7. Which type of reaction occurs when a high-energy particle collides with the nucleus of an atom, converting that atom to an atom of a different element?
   1. addition  
   2. neutralization  
   3. substitution  
   4. transmutation

8. Which particle is emitted when an atom of $^{85}$Kr spontaneously decays?
   1. an alpha particle  
   2. a beta particle  
   3. a neutron  
   4. a proton

9. What is a problem commonly associated with nuclear power facilities?
   1. A small quantity of energy is produced.  
   2. Reaction products contribute to acid rain.  
   3. It is impossible to control nuclear fission.  
   4. It is difficult to dispose of wastes.

10. Which particle has the greatest mass?
    1. an alpha particle  
    2. a beta particle  
    3. a neutron  
    4. a positron

11. A beta particle may be spontaneously emitted from
    1. a ground-state electron  
    2. a stable nucleus  
    3. an excited electron  
    4. an unstable nucleus

12. Which nuclide is used to investigate human thyroid gland disorders?
    1. carbon-14  
    2. potassium-37  
    3. cobalt-60  
    4. iodine-131

13. Which nuclear emission has no charge and no mass?
    1. alpha particle  
    2. beta particle  
    3. gamma ray  
    4. positron
14. Which radioisotope is matched with its decay mode?

(1) H-3 and γ  (3) N-16 and α
(2) K-42 and β⁻  (4) P-32 and β⁻

15. Which reaction is accomplished by the release of the greatest amount of energy?

1. combustion of 10. g of propane  3. nuclear fission of 10. g of uranium
2. electrolysis of 10. g of water  4. oxidation of 10. g of iron

16. Which nuclides are used to date the remains of a once-living organism?

1. C-14 and C₁₂  3. I-131 and Xe-131
2. Co-60 and Co-59  4. U-238 and Pb-206

17. Energy is released during the fission of Pu-239 atoms as a result of the

1. formation of covalent bonds  3. conversion of matter to energy
2. formation of ionic bonds  4. conversion of energy to matter

18. Atoms of I-131 spontaneously decay when the

1. stable nuclei emit alpha particles  3. unstable nuclei emit alpha particles
2. stable nuclei emit beta particles  4. unstable nuclei emit beta particles

19. Which type of reaction releases the greatest amount of energy per mole of reactant?

1. combustion  3. nuclear fusion
2. decomposition  4. oxidation-reduction

20. The spontaneous decay of an atom is called

1. ionization  3. combustion
2. crystallization  4. transmutation

21. Which type of emission has the highest penetrating power?

1. alpha  3. positron
2. beta  4. gamma

22. Which of these types of nuclear radiation has the greatest penetrating power?

1. alpha  3. neutron
2. beta  4. gamma

23. Radioactive cobalt-60 is used in radiation therapy treatment. Cobalt-60 undergoes beta decay. This type of nuclear reaction is called

1. natural transmutation  3. nuclear fusion
2. artificial transmutation  4. nuclear fission
24. The accompanying chart shows the spontaneous nuclear decay of U-238 to Th-234 to Pa-234 to U-234. What is the correct order of nuclear decay modes for the change from U-238 to U-234?

(1) $\beta^-$ decay, $\gamma$ decay, $\beta^-$ decay
(2) $\beta^-$ decay, $\beta^+$ decay, $\alpha$ decay
(3) $\alpha$ decay, $\alpha$ decay, $\beta^-$ decay
(4) $\alpha$ decay, $\beta^-$ decay, $\beta^-$ decay

25. Atoms of one element are converted to atoms of another element through

1. fermentation
2. oxidation
3. polymerization
4. transmutation

26. What is the decay mode of $^{37}$K?

(1) $\beta^-$
(2) $\beta^+$
(3) $\gamma$
(4) $\alpha$

27. Which nuclear emission has the greatest penetrating power?

1. alpha particle
2. beta particle
3. gamma radiation
4. positron

28. What is the mass number of an alpha particle?

1. 1
2. 2
3. 0
4. 4

29. A nuclear reaction in which two light nuclei combine to form a more massive nucleus is called

1. addition
2. fission
3. fusion
4. substitution
30. The nucleus of a radium-226 atom is unstable, which causes the nucleus to spontaneously
1. absorb electrons 3. decay
2. absorb protons 4. oxidize

31. A serious risk factor associated with the operation of a nuclear power plant is the production of
1. acid rain 3. greenhouse gases, such as CO₂
2. helium gas 4. radioisotopes with long half-lives

32. Which nuclide is paired with a specific use of that nuclide?
1. carbon-14, treatment of cancer
2. cobalt-60, dating of rock formations
3. iodine-131, treatment of thyroid disorders
4. uranium-238, dating of once-living organisms

33. A change in the nucleus of an atom that converts the atom from one element to another element is
1. combustion
2. neutralization
3. polymerization
4. transmutation

34. Which particle is emitted from a hydrogen-3 nucleus when it undergoes radioactive decay?
(1) α  (3) β⁺
(2) β⁻  (4) γ

35. In which type of reaction is an atom of one element converted to an atom of a different element?
1. decomposition
2. neutralization
3. saponification
4. transmutation

36. Which isotope is used to treat cancer?
1. C-14
2. U-238
3. Co-60
4. Pb-206

37. Base your answer on the information below.
In living organisms, the ratio of the naturally occurring isotopes of carbon, C-12 to C-13 to C-14, is fairly consistent. When an organism such as a woolly mammoth died, it stopped taking in carbon, and the amount of C-14 present in the mammoth began to decrease. For example, one fossil of a woolly mammoth is found to have 1/32 of the amount of C-14 found in a living organism. Identify the type of nuclear reaction that caused the amount of C-14 in the woolly mammoth to decrease after the organism died. [1]

38. a State one possible advantage of using nuclear power instead of burning fossil fuels. [1]
b State one possible risk of using nuclear power. [1]
c If animals feed on plants that have taken up Sr-90, the Sr-90 can find its way into their bone structure. Explain one danger to the animals. [1]
39. Base your answer on the information below and on your knowledge of chemistry.

**Nuclear Waste Storage Plan for Yucca Mountain**

In 1978, the U.S. Department of Energy began a study of Yucca Mountain which is located 90 miles from Las Vegas, Nevada. The study was to determine if Yucca Mountain would be suitable for a long-term burial site for high-level radioactive waste. A three-dimensional (3-D) computer scale model of the site was used to simulate the Yucca Mountain area. The computer model study for Yucca Mountain included such variables as: the possibility of earthquakes, predicted water flow through the mountain, increased rainfall due to climate changes, radioactive leakage from the waste containers, and increased temperatures from the buried waste within the containers. The containers that will be used to store the radioactive waste are designed to last 10,000 years. Within the 10,000-year time period, cesium and strontium, the most powerful radioactive emitters, would have decayed. Other isotopes found in the waste would decay more slowly, but are not powerful radioactive emitters. In 1998, scientists discovered that the compressed volcanic ash making up Yucca Mountain was full of cracks. Because of the arid climate, scientists assumed that rainwater would move through the cracks at a slow rate. However, when radioactive chlorine-36 was found in rock samples at levels halfway through the mountain, it was clear that rainwater had moved quickly down through Yucca Mountain. It was only 50 years earlier when this chlorine-36 isotope had contaminated rainwater during atmospheric testing of the atom bomb. Some opponents of the Yucca Mountain plan believe that the uncertainties related to the many variables of the computer model result in limited reliability of its predictions. However, advocates of the plan believe it is safer to replace the numerous existing radioactive burial sites around the United States with the one site at Yucca Mountain. Other opponents of the plan believe that transporting the radioactive waste to Yucca Mountain from the existing 131 burial sites creates too much danger to the United States. In 2002, after many years of political debate, a final legislative vote approved the development of Yucca Mountain to replace the existing 131 burial sites.

a) Scientists assume that a manufacturing defect would cause at least one of the waste containers stored in the Yucca Mountain repository to leak within the first 1,000 years. State one possible effect such a leak could have on the environment near Yucca Mountain. [1]

b) State one risk associated with leaving radioactive waste in the 131 sites around the country where it is presently stored. [1]

c) The information states "Within the 10,000-year time period, cesium and strontium, the most powerful radioactive emitters, would have decayed." Use information from Reference Table N to support this statement. [1]

d) Why is water flow a crucial factor in deciding whether Yucca Mountain is a suitable burial site? [1]

40. Base your answer on the information below.

A battery-operated smoke detector produces an alarming sound when its electrical sensor detects smoke particles. Some ionizing smoke detectors contain the radioisotope americium-241, which undergoes alpha decay and has a half-life of 433 years. The emitted alpha particles ionize gas molecules in the air. As a result, an electric current flows through the detector. When smoke particles enter the detector, the flow of ions is interrupted, causing the alarm to sound.

a) State one scientific reason why Am-241 is a more appropriate radioactive source than Fr-220 in an ionizing smoke detector. [1]

b) Explain, in terms of particle behavior, why smoke particles cause the detector alarm to sound. [1]
41. Polonium-210 occurs naturally, but is scarce. Polonium-210 is primarily used in devices designed to eliminate static electricity in machinery. It is also used in brushes to remove dust from camera lenses. Polonium-210 can be created in the laboratory by bombarding bismuth-209 with neutrons to create bismuth-210. The bismuth-210 undergoes beta decay to produce polonium-210. Polonium-210 has a half-life of 138 days and undergoes alpha decay. State one beneficial use of Po-210. [1]

42. In the 1920s, paint used to inscribe the numbers on watch dials was composed of a luminescent (glow-in-the-dark) mixture. The powdered-paint base was a mixture of radium salts and zinc sulfide. As the paint was mixed, the powdered base became airborne and drifted throughout the workroom causing the contents of the workroom, including the painters' clothes and bodies, to glow in the dark. The paint is luminescent because radiation from the radium salts strikes a scintillator. A scintillator is a material that emits visible light in response to ionizing radiation. In watchdial paint, zinc sulfide acts as the scintillator. Radium present in the radium salts decomposes spontaneously, emitting alpha particles. These particles can cause damage to the body when they enter human tissue. Alpha particles are especially harmful to the blood, liver, lungs, and spleen because they can alter genetic information in the cells. Radium can be deposited in the bones because it substitutes for calcium.
   a) Write the notation for the alpha particles emitted by radium in the radium salts. [1]
   b) How can particles emitted from radioactive nuclei damage human tissue? [1]

43. Base your answer on the reading passage shown and on your knowledge of chemistry.

   A Glow in the Dark, and Scientific Peril
   The [Marie and Pierre] Curies set out to study radioactivity in 1898. Their first accomplishment was to show that radioactivity was a property of atoms themselves. Scientifically, that was the most important of their findings, because it helped other researchers refine their understanding of atomic structure.
   More famous was their discovery of polonium and radium. Radium was the most radioactive substance the Curies had encountered. Its radioactivity is due to the large size of the atom, which makes the nucleus unstable and prone to decay, usually to radon and then lead, by emitting particles and energy as it seeks a more stable configuration.
   Marie Curie struggled to purify radium for medical uses, including early radiation treatment for tumors. But radium's bluish glow caught people's fancy, and companies in the United States began mining it and selling it as a novelty: for glow-in-the-dark light pulls, for instance, and bogus cure-all patent medicines that actually killed people.
   What makes radium so dangerous is that it forms chemical bonds in the same way as calcium, and the body can mistake it for calcium and absorb it into the bones. Then, it can bombard cells with radiation at close range, which may cause bone tumors or bone-marrow damage that can give rise to anemia or leukemia.
   State one risk associated with the use of radium. [1]

44. The radioisotopes carbon-14 and nitrogen-16 are present in a living organism. Carbon-14 is commonly used to date a once-living organism. Explain why N-16 is a poor choice for radioactive dating of a bone.
45. Base your answer on the information below.
A U-238 atom decays to a Pb-206 atom through a series of steps. Each point on the accompanying graph represents a nuclide and each arrow represents a nuclear decay mode.
a) Based on this graph, what particle is emitted during the nuclear decay of a Po-218 atom? [1]
b) Explain why the U-238 disintegration series ends with the nuclide Pb-206. [1]

46. Base your answer on the information below.
Scientists are investigating the production of energy using hydrogen-2 nuclei (deuterons) and hydrogen-3 nuclei (tritons). The balanced equation in the accompanying image represents one nuclear reaction between two deuterons.
State, in terms of subatomic particles, how a deuteron differs from a triton. [1]

\[ ^2_1H + ^2_1H \rightarrow ^3_2He + ^1_0n + 5.23 \times 10^{-13} \text{ J} \]

**Natural and artificial transmutation:**

1. Which balanced equation represents natural transmutation?

   (1) \[ ^9_4Be + ^1_1H \rightarrow ^6_3Li + ^4_2He \]
   (2) \[ ^{14}_7N + ^2_4He \rightarrow ^{17}_8O + ^1_1H \]
   (3) \[ ^{239}_94Pu + ^1_0n \rightarrow ^{244}_{94}Ce + ^{94}_{38}Kr + ^2_1He \]
   (4) \[ ^{239}_{92}U \rightarrow ^{234}_{90}Th + ^4_2He \]

2. Which nuclear equation represents a natural transmutation?

   (1) \[ ^9_4Be + ^1_1H \rightarrow ^6_3Li + ^4_2He \]
   (2) \[ ^{27}_{13}Al + ^3_2He \rightarrow ^{30}_{15}P + ^1_0n \]
   (3) \[ ^{14}_7N + ^2_4He \rightarrow ^{17}_8O + ^1_1H \]
   (4) \[ ^{235}_{92}U \rightarrow ^{231}_{89}Th + ^4_2He \]

3. Given the nuclear equation: \( ^{58}_{29}\text{Cu} \rightarrow ^{58}_{28}\text{Ni} + X \)

   What nuclear particle is represented by \( X \)? [1]
4. Complete the nuclear equation shown. Include the symbol, atomic number, and mass number for the missing particle. [1]

\[ ^{40}_{19}K \rightarrow \_\_e + \_\_\_ \]

5. Base your answer on the information below.
A battery-operated smoke detector produces an alarming sound when its electrical sensor detects smoke particles. Some ionizing smoke detectors contain the radioisotope americium-241, which undergoes alpha decay and has a half-life of 433 years. The emitted alpha particles ionize gas molecules in the air. As a result, an electric current flows through the detector. When smoke particles enter the detector, the flow of ions is interrupted, causing the alarm to sound.
Complete the nuclear equation on the answer sheet or on a separate piece of paper for the decay of Am-241. Your response must include the symbol, mass number, and atomic number of each product. [2]

\[ ^{241}_{95}Am \rightarrow \_\_\_ + \_\_\_ \]

6. Base your answer on the information below.
Polonium-210 occurs naturally, but is scarce. Polonium-210 is primarily used in devices designed to eliminate static electricity in machinery. It is also used in brushes to remove dust from camera lenses. Polonium-210 can be created in the laboratory by bombarding bismuth-209 with neutrons to create bismuth-210. The bismuth-210 undergoes beta decay to produce polonium-210. Polonium-210 has a half-life of 138 days and undergoes alpha decay.
Complete the nuclear equation on the answer sheet or on a separate piece of paper for the decay of Po-210, by writing a notation for the missing product. [1]

7. Base your answer on the reading passage shown and on your knowledge of chemistry.

**A Glow in the Dark, and Scientific Peril**
The [Marie and Pierre] Curies set out to study radioactivity in 1898. Their first accomplishment was to show that radioactivity was a property of atoms themselves. Scientifically, that was the most important of their findings, because it helped other researchers refine their understanding of atomic structure.
More famous was their discovery of polonium and radium. Radium was the most radioactive substance the Curies had encountered. Its radioactivity is due to the large size of the atom, which makes the nucleus unstable and prone to decay, usually to radon and then lead, by emitting particles and energy as it seeks a more stable configuration.
Marie Curie struggled to purify radium for medical uses, including early radiation treatment for tumors. But radium's bluish glow caught people's fancy, and companies in the United States began mining it and selling it as a novelty: for glow-in-the-dark light pulls, for instance, and bogus cure-all patent medicines that actually killed people.
What makes radium so dangerous is that it forms chemical bonds in the same way as calcium, and the body can mistake it for calcium and absorb it into the bones. Then, it can bombard cells with radiation at close range, which may cause bone tumors or bone-marrow damage that can give rise to anemia or leukemia.


Using Reference Table \( N \), complete the equation provided on a separate piece of paper for the nuclear decay of \( ^{226}_{88} \text{Ra} \). Include both atomic number and mass number for each particle. [1]
8. Base your answer to the question on the information below.
The radioisotopes carbon-14 and nitrogen-16 are present in a living organism. Carbon-14 is commonly used to date a once-living organism.
Complete the accompanying nuclear equation in the space provided or on a separate piece of paper for the decay of C-14. Include both the atomic number and the mass number of the missing particle.
\[ ^{14}_{6}C \rightarrow \quad \quad +^{0}_{-1}e \]

9. Base your answer on the information below.
The fossilized remains of a plant were found at a construction site. The fossilized remains contain 1/16 the amount of carbon-14 that is present in a living plant.
Complete the nuclear equation on the answer sheet, or, if taken online, on a separate piece of paper, for the decay of C-14. Your response must include the atomic number, the mass number, and the symbol of the missing particle. [1]
\[ ^{14}_{6}C \rightarrow -^{0}_{1}e + \quad \quad \quad \]

10. Base your answer on the information below.
The radioisotope uranium-238 occurs naturally in Earth's crust. The disintegration of this radioisotope is the first in a series of spontaneous decays.
The sixth decay in this series produces the radioisotope radon-222. The decay of radon-222 produces the radioisotope polonium-218 that has a half life of 3.04 minutes. Eventually, the stable isotope lead-206 is produced by the alpha decay of an unstable nuclide.
Complete the nuclear equation shown on the accompanying diagram for the decay of the unstable nuclide that produces Pb-206, by writing a notation for the missing nuclide. [1]
\[ \quad \quad \quad \rightarrow ^{4}_{2}He + ^{206}_{82}Pb \]

**Fission and fusion:**

1. In a nuclear fusion reaction, the mass of the products is
   1. less than the mass of the reactants because some of the mass has been converted to energy
   2. less than the mass of the reactants because some of the energy has been converted to mass
   3. more than the mass of the reactants because some of the mass has been converted to energy
   4. more than the mass of the reactants because some of the energy has been converted to mass

2. Which balanced equation represents nuclear fusion?
   \( (1) \ \ ^{0}_{1}n + ^{238}_{92}U \rightarrow ^{142}_{56}Ba + ^{89}_{36}Kr + 3^{0}_{1}n \)
   \( (2) \ \ ^{226}_{88}Ra \rightarrow ^{222}_{86}Rn + ^{4}_{2}He \)
   \( (3) \ \ ^{6}_{3}Li + ^{1}_{0}n \rightarrow ^{3}_{1}H + ^{3}_{1}He \)
   \( (4) \ \ ^{2}_{1}H + ^{3}_{1}H \rightarrow ^{4}_{2}He + ^{0}_{1}n \)
3. A nuclear fission reaction and a nuclear fusion reaction are similar because both reactions
1. form heavy nuclides from light nuclides 2. release a large amount of energy
2. form light nuclides from heavy nuclides 3. absorb a large amount of energy

4. Given the balanced equation representing a nuclear reaction:
Which particle is represented by X?

\[ {^{235}}_{92}\text{U} + {^1}_0\text{n} \rightarrow {^{142}}_{56}\text{Ba} + {^{91}}_{37}\text{Kr} + 3X + \text{energy} \]

(1) \( ^0\text{e} \)  (3) \( \frac{1}{2}\text{He} \)
(2) \( ^1\text{H} \)  (4) \( ^0\text{n} \)

5. When a uranium-235 nucleus absorbs a slow-moving neutron, different nuclear reactions may occur. One of these possible reactions is represented by the complete, balanced Equation 1.
For this reaction, the sum of the masses of the products is slightly less than the sum of the masses of the reactants. Another possible reaction of U-235 is represented by the incomplete, balanced Equation 2.
a) Identify the type of nuclear reaction represented by equation 1. [1]
b) Write a notation for the missing product in equation 2. [1]

Equation 1: \( {^{235}}_{92}\text{U} + {^1}_0\text{n} \rightarrow {^{92}}_{36}\text{Kr} + {^{142}}_{56}\text{Ba} + 2{^1}_0\text{n} + \text{energy} \)

Equation 2: \( {^{235}}_{92}\text{U} + {^1}_0\text{n} \rightarrow {^{92}}_{38}\text{Sr} + _____ + 2{^1}_0\text{n} + \text{energy} \)

6. Scientists are investigating the production of energy using hydrogen-2 nuclei (deuterons) and hydrogen-3 nuclei (tritons). The balanced equation in the accompanying image represents one nuclear reaction between two deuterons.
Identify the type of nuclear reaction represented by the equation. [1]

\( ^2\text{H} + ^2\text{H} \rightarrow ^3\text{He} + ^1\text{n} + 5.23 \times 10^{-13} \text{ J} \)

**Half life:**

1. The half-life of a radioactive substance is 2.5 minutes. What fraction of the original radioactive substance remains after 10 minutes?
   (1) \( \frac{1}{2} \)  (3) \( \frac{1}{8} \)
   (2) \( \frac{1}{4} \)  (4) \( \frac{1}{16} \)

2. As a sample of the radioactive isotope \(^{131}\text{I}\) decays, its half-life
   1. decreases  3. remains the same
   2. increases
3. Exactly how much time must elapse before 16 grams of potassium-42 decays, leaving 2 grams of the original isotope?
   1. $8 \times 12.4$ hours
   2. $2 \times 12.4$ hours
   3. $3 \times 12.4$ hours
   4. $4 \times 12.4$ hours

4. Based on Reference Table $N$, what fraction of a sample of gold-198 remains radioactive after 2.69 days?
   1. $1/4$
   2. $1/2$
   3. $3/4$
   4. $7/8$

5. According to Reference Table $N$, which pair of isotopes spontaneously decays?
   1. C-12 and N-14
   2. C-12 and N-16
   3. C-14 and N-14
   4. C-14 and N-16

6. What is the half-life of sodium-25 if 1.00 gram of a 16.00-gram sample of sodium-25 remains unchanged after 237 seconds?
   1. 47.4 s
   2. 59.3 s
   3. 79.0 s
   4. 118 s

7. An original sample of K-40 has a mass of 25.00 grams. After $3.9 \times 10^6$ years, 3.125 grams of the original sample remains unchanged. What is the half-life of K-40?
   1. $1.3 \times 10^9$ y
   2. $2.6 \times 10^9$ y
   3. $3.9 \times 10^9$ y
   4. $1.2 \times 10^{10}$ y

8. According to Reference Table $N$, which radioactive isotope will retain only one-eighth (1/8;) its original radioactive atoms after approximately 43 days?
   1. gold-198
   2. iodine-131
   3. phosphorus-32
   4. radon-222

9. An original sample of the radioisotope fluorine-21 had a mass of 80.0 milligrams. Only 20.0 milligrams of this original sample remain unchanged after 8.32 seconds. What is the half-life of fluorine-21?
   1. 1.04 s
   2. 2.08 s
   3. 4.16 s
   4. 8.32 s

10. Approximately what fraction of an original Co-60 sample remains after 21 years?
    1. $1/2$
    2. $1/4$
    3. $1/8$
    4. $1/16$

11. What is the half-life of a radioisotope if 25.0 grams of an original 200-gram sample of the isotope remains unchanged after 11.46 days?
    1. 2.87 d
    2. 3.82 d
    3. 11.46 d
    4. 34.38 d
12. Which nuclide is listed with its half-life and decay mode?

(1) K-37, 1.24 h, α
(2) N-16, 7.2 s, β⁻
(3) Rn-222, 1.6 × 10³ y, α
(4) U-235, 7.1 × 10⁸ y, β⁻

13. Base your answer on the information below and on your knowledge of chemistry.

**Nuclear Waste Storage Plan for Yucca Mountain**

In 1978, the U.S. Department of Energy began a study of Yucca Mountain which is located 90 miles from Las Vegas, Nevada. The study was to determine if Yucca Mountain would be suitable for a long-term burial site for high-level radioactive waste. A three-dimensional (3-D) computer scale model of the site was used to simulate the Yucca Mountain area. The computer model study for Yucca Mountain included such variables as: the possibility of earthquakes, predicted water flow through the mountain, increased rainfall due to climate changes, radioactive leakage from the waste containers, and increased temperatures from the buried waste within the containers. The containers that will be used to store the radioactive waste are designed to last 10,000 years. Within the 10,000-year time period, cesium and strontium, the most powerful radioactive emitters, would have decayed. Other isotopes found in the waste would decay more slowly, but are not powerful radioactive emitters. In 1998, scientists discovered that the compressed volcanic ash making up Yucca Mountain was full of cracks. Because of the arid climate, scientists assumed that rainwater would move through the cracks at a slow rate. However, when radioactive chlorine-36 was found in rock samples at levels halfway through the mountain, it was clear that rainwater had moved quickly down through Yucca Mountain. It was only 50 years earlier when this chlorine-36 isotope had contaminated rainwater during atmospheric testing of the atom bomb. Some opponents of the Yucca Mountain plan believe that the uncertainties related to the many variables of the computer model result in limited reliability of its predictions. However, advocates of the plan believe it is safer to replace the numerous existing radioactive burial sites around the United States with the one site at Yucca Mountain. Other opponents of the plan believe that transporting the radioactive waste to Yucca Mountain from the existing 131 burial sites creates too much danger to the United States. In 2002, after many years of political debate, a final legislative vote approved the development of Yucca Mountain to replace the existing 131 burial sites.

If a sample of cesium-137 is stored in a waste container in Yucca Mountain, how much time must elapse until only 1/32 of the original sample remains unchanged? [1]

14. Base your answer on the information below.

In living organisms, the ratio of the naturally occurring isotopes of carbon, C-12 to C-13 to C-14, is fairly consistent. When an organism such as a woolly mammoth died, it stopped taking in carbon, and the amount of C-14 present in the mammoth began to decrease. For example, one fossil of a woolly mammoth is found to have 1/32 of the amount of C-14 found in a living organism. Determine the total time that has elapsed since this woolly mammoth died. [1]

15. Base your answer on the information below and the accompanying equations.

When a uranium-235 nucleus absorbs a slow-moving neutron, different nuclear reactions may occur. One of these possible reactions is represented by the complete, balanced Equation 1. For this reaction, the sum of the masses of the products is slightly less than the sum of the masses of the reactants. Another possible reaction of U-235 is represented by the incomplete, balanced Equation 2.
Determine the half-life of krypton-92 if only 6.0 milligrams of an original 96.0-milligram sample remains unchanged after 7.36 seconds. [1]

\[
\text{Equation 1: } ^{235}\text{U} + ^{0}\text{n} \rightarrow ^{92}\text{Kr} + ^{142}\text{Ba} + 2^1\text{n} + \text{energy}
\]

\[
\text{Equation 2: } ^{235}\text{U} + ^{0}\text{n} \rightarrow ^{92}\text{Sr} + \text{______} + 2^1\text{n} + \text{energy}
\]

16. Base your answer on the information below.
Polonium-210 occurs naturally, but is scarce. Polonium-210 is primarily used in devices designed to eliminate static electricity in machinery. It is also used in brushes to remove dust from camera lenses. Polonium-210 can be created in the laboratory by bombarding bismuth-209 with neutrons to create bismuth-210. The bismuth-210 undergoes beta decay to produce polonium-210. Polonium-210 has a half-life of 138 days and undergoes alpha decay. Determine the total mass of an original 28.0-milligram sample of Po-210 that remains unchanged after 414 days. [1]

17. Base your answer on the reading passage shown and on your knowledge of chemistry.

A Glow in the Dark, and Scientific Peril

The [Marie and Pierre] Curies set out to study radioactivity in 1898. Their first accomplishment was to show that radioactivity was a property of atoms themselves. Scientifically, that was the most important of their findings, because it helped other researchers refine their understanding of atomic structure.

More famous was their discovery of polonium and radium. Radium was the most radioactive substance the Curies had encountered. Its radioactivity is due to the large size of the atom, which makes the nucleus unstable and prone to decay, usually to radon and then lead, by emitting particles and energy as it seeks a more stable configuration.

Marie Curie struggled to purify radium for medical uses, including early radiation treatment for tumors. But radium's bluish glow caught people's fancy, and companies in the United States began mining it and selling it as a novelty: for glow-in-the-dark light pulls, for instance, and bogus cure-all patent medicines that actually killed people.

What makes radium so dangerous is that it forms chemical bonds in the same way as calcium, and the body can mistake it for calcium and absorb it into the bones. Then, it can bombard cells with radiation at close range, which may cause bone tumors or bone-marrow damage that can give rise to anemia or leukemia.


If a scientist purifies 1.0 gram of radium-226, how many years must pass before only 0.50 gram of the original radium-226 sample remains unchanged? [1]

18. Base your answer to the question on the information below.
The radioisotopes carbon-14 and nitrogen-16 are present in a living organism. Carbon-14 is commonly used to date a once-living organism.

A sample of wood is found to contain 1/8 as much C-14 as is present in the wood of a living tree. What is the approximate age, in years, of this sample of wood?
The fossilized remains of a plant were found at a construction site. The fossilized remains contain 1/16 the amount of carbon-14 that is present in a living plant.
Determine the approximate age of these fossilized remains. [1]

20. Base your answer on the information below.
The radioisotope uranium-238 occurs naturally in Earth's crust. The disintegration of this radioisotope is the first in a series of spontaneous decays.
The sixth decay in this series produces the radioisotope radon-222. The decay of radon-222 produces the radioisotope polonium-218 that has a half life of 3.04 minutes. Eventually, the stable isotope lead-206 is produced by the alpha decay of an unstable nuclide.
Determine the original mass of a sample of Po-218, if 0.50 milligram of the sample remains unchanged after 12.16 minutes. [1]

Math skills, lab skills and graphing:

1. A sample of water is being heated from 20°C to 30°C, and the temperature is recorded every 2 minutes. Which table would be most appropriate for recording the data?

<table>
<thead>
<tr>
<th>Time (min)</th>
<th>Temp (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>20</td>
</tr>
<tr>
<td>2</td>
<td>22</td>
</tr>
<tr>
<td>4</td>
<td>24</td>
</tr>
<tr>
<td>6</td>
<td>26</td>
</tr>
<tr>
<td>8</td>
<td>28</td>
</tr>
<tr>
<td>10</td>
<td>30</td>
</tr>
</tbody>
</table>

2. The diagram shown represents a portion of a 100-milliliter graduated cylinder. What is the reading of the meniscus?

1. 35.0 mL
2. 36.0 mL
3. 44.0 mL
4. 45.0 mL

3. Base your answer on the data table, which shows the solubility of a solid solute.
a. On the grid provided (print this page), mark an appropriate scale on the axis labeled "Solute per 100 g of H₂O(g)." An appropriate scale is one that allows a trend to be seen. [1]
b. Plot the data from the data table. Circle and connect the points. [1]
c. Based on the data table, if 15 grams of solute is dissolved in 100 grams of water at 40°C, how
many *more* grams of solute can be dissolved in this solution to make it saturated at 40°C? [1]

### The Solubility of the Solute at Various Temperatures

<table>
<thead>
<tr>
<th>Temperature (°C)</th>
<th>Solute per 100 g of H₂O(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>18</td>
</tr>
<tr>
<td>20</td>
<td>20</td>
</tr>
<tr>
<td>40</td>
<td>24</td>
</tr>
<tr>
<td>60</td>
<td>29</td>
</tr>
<tr>
<td>80</td>
<td>36</td>
</tr>
<tr>
<td>100</td>
<td>49</td>
</tr>
</tbody>
</table>

**Example:**

![Example Image]
4. Base your answer on the information below.
A weather balloon has a volume of 52.5 liters at a temperature of 295 K. The balloon is released and rises to an altitude where the temperature is 252 K.
What pressure, in atmospheres (atm), is equal to 45.6 kPa? [1]

5. Base your answer on the information in the accompanying graph, which relates the numbers of neutrons and protons for specific nuclides of C, N, Ne, and S.
a) Using the point plotted on the graph for neon, complete the accompanying table.
b) Using the point plotted on the graph for nitrogen, what is the neutron-to-proton ratio of this nuclide?

![Graph](image)

**Answer Sheet**

<table>
<thead>
<tr>
<th>Element</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
<th>Mass Number</th>
<th>Nuclide</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>6</td>
<td>6</td>
<td>12</td>
<td>C-12</td>
</tr>
<tr>
<td>N</td>
<td>7</td>
<td>9</td>
<td>16</td>
<td>N-16</td>
</tr>
<tr>
<td>Ne</td>
<td>10</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>S</td>
<td>16</td>
<td>16</td>
<td>32</td>
<td>S-32</td>
</tr>
</tbody>
</table>

6. Base your answer on the information below and on your knowledge of chemistry.
**Nuclear Waste Storage Plan for Yucca Mountain**
In 1978, the U.S. Department of Energy began a study of Yucca Mountain which is located 90 miles from Las Vegas, Nevada. The study was to determine if Yucca Mountain would be suitable for a long-term burial site for high-level radioactive waste. A three-dimensional (3-D) computer scale model of the site was used to simulate the Yucca Mountain area. The computer model study for Yucca Mountain included such variables as: the possibility of earthquakes, predicted water flow through the mountain, increased rainfall due to climate changes, radioactive leakage from the waste
containers, and increased temperatures from the buried waste within the containers. The containers that will be used to store the radioactive waste are designed to last 10,000 years. Within the 10,000-year time period, cesium and strontium, the most powerful radioactive emitters, would have decayed. Other isotopes found in the waste would decay more slowly, but are not powerful radioactive emitters. In 1998, scientists discovered that the compressed volcanic ash making up Yucca Mountain was full of cracks. Because of the arid climate, scientists assumed that rainwater would move through the cracks at a slow rate. However, when radioactive chlorine-36 was found in rock samples at levels halfway through the mountain, it was clear that rainwater had moved quickly down through Yucca Mountain. It was only 50 years earlier when this chlorine-36 isotope had contaminated rainwater during atmospheric testing of the atom bomb. Some opponents of the Yucca Mountain plan believe that the uncertainties related to the many variables of the computer model result in limited reliability of its predictions. However, advocates of the plan believe it is safer to replace the numerous existing radioactive burial sites around the United States with the one site at Yucca Mountain. Other opponents of the plan believe that transporting the radioactive waste to Yucca Mountain from the existing 131 burial sites creates too much danger to the United States. In 2002, after many years of political debate, a final legislative vote approved the development of Yucca Mountain to replace the existing 131 burial sites.

State one uncertainty in the computer model that limits the reliability of this computer model. [1]

7. Base your answer on the accompanying table, which shows the electronegativity of selected elements of Period 2 of the Periodic Table.
   a On the grid shown, or on a piece of graph paper, set up a scale for electronegativity on the y-axis. Plot the data by drawing a best-fit line. [2]
   b Using the graph, predict the electronegativity of nitrogen. [1]
   c For these elements, state the trend in electronegativity in terms of atomic number. [1]
8. In one trial of an investigation, 50.0 milliliters of HCl(aq) of an unknown concentration is titrated with 0.10 M NaOH(aq). During the titration, the total volume of NaOH(aq) added and the corresponding pH value of the reaction mixture are measured and recorded in the accompanying table.

a) Determine the total volume of NaOH(aq) added when the reaction mixture has a pH value of 7.0.
b) On the grid provided, plot the data from the table. Circle and connect the points.

<table>
<thead>
<tr>
<th>Total Volume of NaOH(aq) Added (mL)</th>
<th>pH Value of Reaction Mixture</th>
</tr>
</thead>
<tbody>
<tr>
<td>10.0</td>
<td>1.6</td>
</tr>
<tr>
<td>20.0</td>
<td>2.2</td>
</tr>
<tr>
<td>24.0</td>
<td>2.9</td>
</tr>
<tr>
<td>24.9</td>
<td>3.9</td>
</tr>
<tr>
<td>25.1</td>
<td>10.1</td>
</tr>
<tr>
<td>26.0</td>
<td>11.1</td>
</tr>
<tr>
<td>30.0</td>
<td>11.8</td>
</tr>
</tbody>
</table>

9. Base your answer on the information below and the accompanying table. Bond energy is the amount of energy required to break a chemical bond. The table gives a formula and the carbon-nitrogen bond energy for selected nitrogen compounds. State the relationship between the number of electrons in a carbon-nitrogen bond and carbon-nitrogen bond energy. [1]

<table>
<thead>
<tr>
<th>Selected Nitrogen Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>Compound</td>
</tr>
<tr>
<td>hydrogen cyanide</td>
</tr>
<tr>
<td>isocyanic acid</td>
</tr>
<tr>
<td>methanamine</td>
</tr>
</tbody>
</table>
10. Base your answer on the information below.
Gasoline is a mixture composed primarily of hydrocarbons such as isooctane, which is also known as 2,2,4-trimethylpentane.
Gasoline is assigned a number called an octane rating. Gasoline with an octane rating of 87 performs the same as a mixture that consists of 87% isooctane and 13% heptane.
An alternative fuel, E-85, can be used in some automobiles. This fuel is a mixture of 85% ethanol and 15% gasoline.
State the octane rating of a gasoline sample that performs the same as a mixture consisting of 92% isooctane and 8% heptane. [1]

11. Base your answers on the information below.
In a laboratory, a glass tube is filled with hydrogen gas at a very low pressure. When a scientist applies high voltage between metal electrodes in the tube, light is emitted. The scientist analyzes the light with a spectroscope and observes four distinct spectral lines. The accompanying table gives the color, frequency, and energy for each of the four spectral lines. The unit for frequency is hertz, Hz.
a. On the grid on the answer sheet or on a separate piece of paper, plot the data from the data table for frequency and energy. Circle and connect the points, including the point (0,0) that has already been plotted and circled for you. [1]
b. A spectral line in the infrared region of the spectrum of hydrogen has a frequency of 2.3 \times 10^{14} \text{ hertz}. Using your graph, estimate the energy associated with this spectral line. [1]

<table>
<thead>
<tr>
<th>Color</th>
<th>Frequency (\times 10^{14} \text{ Hz})</th>
<th>Energy (\times 10^{-19} \text{ J})</th>
</tr>
</thead>
<tbody>
<tr>
<td>red</td>
<td>4.6</td>
<td>3.0</td>
</tr>
<tr>
<td>blue green</td>
<td>6.2</td>
<td>4.1</td>
</tr>
<tr>
<td>blue</td>
<td>6.9</td>
<td>4.6</td>
</tr>
<tr>
<td>violet</td>
<td>7.3</td>
<td>4.8</td>
</tr>
</tbody>
</table>
12. Base your answer on the information below. A hydrate is a compound that has water molecules within its crystal structure. The formula for the hydrate CuSO₄·5H₂O(s) shows that there are five moles of water for every one mole of CuSO₄(s). When CuSO₄·5H₂O(s) is heated, the water within the crystals is released, as represented by the balanced equation below.

CuSO₄·5H₂O(s) → CuSO₄(s) + 5H₂O(g)

A student first masses an empty crucible (a heatresistant container). The student then masses the crucible containing a sample of CuSO₄·5H₂O(s). The student repeatedly heats and masses the crucible and its contents until the mass is constant. The student's recorded experimental data and calculations are shown in the accompanying diagram.

Explain why the sample in the crucible must be heated until the constant mass is reached. [1]

<table>
<thead>
<tr>
<th>Data and calculation before heating:</th>
</tr>
</thead>
<tbody>
<tr>
<td>mass of CuSO₄·5H₂O(s) and crucible</td>
</tr>
<tr>
<td>mass of crucible</td>
</tr>
<tr>
<td>mass of CuSO₄·5H₂O(s)</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Data and calculation after heating to a constant mass:</th>
</tr>
</thead>
<tbody>
<tr>
<td>mass of CuSO₄(s) and crucible</td>
</tr>
<tr>
<td>mass of crucible</td>
</tr>
<tr>
<td>mass of CuSO₄(s)</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Calculation to determine the mass of water:</th>
</tr>
</thead>
<tbody>
<tr>
<td>mass of CuSO₄·5H₂O(s)</td>
</tr>
<tr>
<td>mass of CuSO₄(s)</td>
</tr>
<tr>
<td>mass of H₂O(g)</td>
</tr>
</tbody>
</table>

13. Base your answer on the information below.

The boiling point of a liquid is the temperature at which the vapor pressure of the liquid is equal to the pressure on the surface of the liquid. The heat of vaporization of ethanol is 838 joules per gram. A sample of ethanol has a mass of 65.0 grams and is boiling at 1.00 atmosphere.

a) Based on Table H, what is the temperature of this sample of ethanol? [1]

b) Calculate the minimum amount of heat required to completely vaporize this sample of ethanol. Your response must include both a correct numerical setup and the calculated result. [2]

14. Base your answer on the information below.

The temperature of a sample of a substance is increased from 20.0°C to 160.0°C as the sample absorbs heat at a constant rate of 15 kilojoules per minute at standard pressure. The accompanying graph represents the relationship between temperature and time as the sample is heated.

Determine the total amount of heat required to completely melt this sample at its melting point. [1]
15. Base your answer on the accompanying table.
State the trend in first ionization energy for the elements in the table as the atomic number increases.

**First Ionization Energy of Selected Elements**

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic Number</th>
<th>First Ionization Energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>lithium</td>
<td>3</td>
<td>520</td>
</tr>
<tr>
<td>sodium</td>
<td>11</td>
<td>496</td>
</tr>
<tr>
<td>potassium</td>
<td>19</td>
<td>419</td>
</tr>
<tr>
<td>rubidium</td>
<td>37</td>
<td>403</td>
</tr>
<tr>
<td>cesium</td>
<td>55</td>
<td>376</td>
</tr>
</tbody>
</table>

16. Base your answer on the information below.
An investigation was conducted to study the effect of the concentration of a reactant on the total time needed to complete a chemical reaction. Four trials of the same reaction were performed. In each trial the initial concentration of the reactant was different. The time needed for the chemical reaction to be completed was measured. The data for each of the four trials are shown in the accompanying table.
a. On the grid shown or on a separate sheet of paper, mark an appropriate scale on the axis labeled "Reaction Time (s)." An appropriate scale is one that allows a trend to be seen. [1]
b. On the same grid, plot the data from the data table. Circle and connect the points as shown in the accompanying diagram. [1]
c. State the effect of the concentration of the reactant on the rate of the chemical reaction. [1]
17. Base your answer on the information below. Using burets, a student titrated a sodium hydroxide solution of unknown concentration with a standard solution of 0.10 M hydrochloric acid. The data are recorded in the accompanying table. Determine both the total volume of HCl(aq) and the total volume of NaOH(aq) used in the titration. [1]

<table>
<thead>
<tr>
<th>Trial</th>
<th>Initial Concentration (M)</th>
<th>Reaction Time (s)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.020</td>
<td>11</td>
</tr>
<tr>
<td>2</td>
<td>0.015</td>
<td>14</td>
</tr>
<tr>
<td>3</td>
<td>0.010</td>
<td>23</td>
</tr>
<tr>
<td>4</td>
<td>0.005</td>
<td>58</td>
</tr>
</tbody>
</table>

**Example:**

**Titration Data**

<table>
<thead>
<tr>
<th>Solution</th>
<th>HCl(aq)</th>
<th>NaOH(aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial Buret Reading (mL)</td>
<td>15.50</td>
<td>5.00</td>
</tr>
<tr>
<td>Final Buret Reading (mL)</td>
<td>25.00</td>
<td>8.80</td>
</tr>
</tbody>
</table>

___________ mL HCl(aq) and __________ mL NaOH(aq)

18. Base your answer on the information below. The accompanying graph shows a compound being cooled at a constant rate starting in the liquid phase at 75°C and ending at 15°C.

a) What is the freezing point of the compound, in degrees Celsius? [1]
b) A different experiment was conducted with another sample of the same compound starting in the solid phase. The sample was heated at a constant rate from 15°C to 75°C. On the graph shown or on a separate sheet of paper, draw the resulting heating curve. [1]
```
The temperature of a sample of a substance is increased from 20°C to 160°C as the sample absorbs
heat at a constant rate of 15 kilojoules per minute at standard pressure. The accompanying graph
represents the relationship between temperature and time as the sample is heated.
What is the total time this sample is in the liquid phase, only? [1]
```

```
20. Base your answer on the accompanying table.
a. On the grid provided or on a separate piece of paper, mark an appropriate scale on the axis labeled
"First Ionization Energy (kJ/mol)." An appropriate scale is one that allows a trend to be seen.
b. On the same grid, plot the data from the table. Circle and connect the points.
```
Electron affinity is defined as the energy released when an atom and an electron react to form a negative ion. The data for Group 1 elements are shown in the table.

On the grid provided, or on a separate sheet of paper, draw a graph to show the relationship between each member of Group 1 and its electron affinity by following the directions below.

- Label the y-axis "Electron Affinity" and choose an appropriate scale. Label the x-axis "Atomic Number" and choose an appropriate scale. [1]

21. Electron affinity is defined as the energy released when an atom and an electron react to form a negative ion. The data for Group 1 elements are shown in the table.

On the grid provided, or on a separate sheet of paper, draw a graph to show the relationship between each member of Group 1 and its electron affinity by following the directions below.

- Label the y-axis "Electron Affinity" and choose an appropriate scale. Label the x-axis "Atomic Number" and choose an appropriate scale. [1]
b Plot the data from the data table and connect the points with straight lines. [1]
c Using your graph, estimate the electron affinity of Rb, in kilojoules/mole. [1]

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic Number</th>
<th>Electron Affinity in kJ/mole</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cs</td>
<td>55</td>
<td>45.5</td>
</tr>
<tr>
<td>H</td>
<td>1</td>
<td>72.8</td>
</tr>
<tr>
<td>K</td>
<td>19</td>
<td>46.4</td>
</tr>
<tr>
<td>Li</td>
<td>3</td>
<td>59.8</td>
</tr>
<tr>
<td>Na</td>
<td>11</td>
<td>52.9</td>
</tr>
<tr>
<td>Rb</td>
<td>37</td>
<td>?</td>
</tr>
</tbody>
</table>

22. The table provided (see image) gives information about two isotopes of element X. Calculate the average atomic mass of element X.
   - Show a correct numerical setup in the space provided on the answer sheet. [1]
   - Record your answer. [1]
Express your answer to the correct number of significant figures. [1]

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Mass</th>
<th>Relative Abundance</th>
</tr>
</thead>
<tbody>
<tr>
<td>X-10</td>
<td>10.01</td>
<td>19.91%</td>
</tr>
<tr>
<td>X-11</td>
<td>11.01</td>
<td>80.09%</td>
</tr>
</tbody>
</table>

23. Base your answer on the electronegativity values and atomic numbers of fluorine, chlorine, bromine, and iodine that are listed on Reference Table S.
   a. On the grid provided (print this page), mark an appropriate scale on the axis labeled "Electronegativity." An appropriate scale is one that allows a trend to be seen. [1]
   b. On the same grid, plot the electronegativity and atomic number data from Reference Table S. Circle and connect the points. [1]
   Note: This question originally appeared on the NYS Regents exam as questions 58 and 59.

24. Base your answer on the information and data table.
   Indigestion may be caused by excess stomach acid (hydrochloric acid). Some products used to treat indigestion contain magnesium hydroxide. The magnesium hydroxide neutralizes some of the stomach acid.
   The amount of acid that can be neutralized by three different brands of antacids is shown in the data table.
   Which antacid brand neutralizes the most acid per gram of antacid tablet? [1]
Base your answers to the question on the properties of propanone. A liquid's boiling point is the temperature at which its vapor pressure is equal to the atmospheric pressure. Using Reference Table \( H \), what is the boiling point (in centigrade) of propanone at an atmospheric pressure of 70 kPa?

26. Base your answer on the information below.
A student performed a laboratory activity to observe the reaction between aluminum foil and an aqueous copper(II) chloride solution. The reaction is represented by the accompanying balanced equation.

\[
2\text{Al}(s) + 3\text{CuCl}_2(aq) \rightarrow 3\text{Cu}(s) + 2\text{AlCl}_3(aq) + \text{energy}
\]

The procedures and corresponding observations for the activities are given in the accompanying table. State one safety procedure the student should perform after completing the laboratory activity. [1]

<table>
<thead>
<tr>
<th>Antacid Brand</th>
<th>Mass of Antacid Tablet (g)</th>
<th>Volume of HCl(aq) Neutralized (mL)</th>
</tr>
</thead>
<tbody>
<tr>
<td>( X )</td>
<td>2.00</td>
<td>25.20</td>
</tr>
<tr>
<td>( Y )</td>
<td>1.20</td>
<td>18.65</td>
</tr>
<tr>
<td>( Z )</td>
<td>1.75</td>
<td>22.50</td>
</tr>
</tbody>
</table>

27. Base your answer on the information below.
At room temperature, a reaction occurs when KIO\(_3\)(aq) is mixed with NaHSO\(_3\)(aq) that contains a small amount of starch. The colorless reaction mixture turns dark blue after a period of time that depends on the concentration of the reactants.

In a laboratory, 12 drops of a 0.02 M NaHSO\(_3\)(aq) solution containing starch were placed in each of six test tubes. A different number of drops of 0.02 M KIO\(_3\)(aq) and enough water to maintain a constant volume were added to each test tube and the time for the dark-blue color to appear was measured. The data were recorded in the accompanying table.

On the accompanying grid or on a separate piece of paper:
- Mark an appropriate scale on the axis labeled "Time (s)." [1]
- Plot the data from the data table. Circle and connect the points. [1]
28. Base your answers on the information below.
The ionic radii of some Group 2 elements are given in the accompanying table.

<table>
<thead>
<tr>
<th>Test Tube</th>
<th>A</th>
<th>B</th>
<th>C</th>
<th>D</th>
<th>E</th>
<th>F</th>
</tr>
</thead>
<tbody>
<tr>
<td>Number of Drops of 0.02 M KIO₃(aq)</td>
<td>2</td>
<td>4</td>
<td>6</td>
<td>8</td>
<td>10</td>
<td>12</td>
</tr>
<tr>
<td>Time for Dark-Blue Color to Appear (s)</td>
<td>210</td>
<td>88</td>
<td>49</td>
<td>39</td>
<td>33</td>
<td>27</td>
</tr>
</tbody>
</table>

a. On the grid on the answer sheet or on a separate piece of paper, mark an appropriate scale on the axis labeled "Ionic Radius (pm)." [1]
b. On the same grid, plot the data from the data table. Circle and connect the points. [1]
c. Estimate the ionic radius of strontium. [1]
d. State the trend in ionic radius as the elements in Group 2 are considered in order of increasing atomic number. [1]
e. Explain, in terms of electrons, why the ionic radius of a Group 2 element is smaller than its atomic radius. [1]
29. Base your answer on the information below.
At 550°C, 1.00 mole of CO₂(g) and 1.00 mole of H₂(g) are placed in a 1.00-liter reaction vessel. The substances react to form CO(g) and H₂O(g). Changes in the concentrations of the reactants and the concentrations of the products are shown in the accompanying graph.
Determine the change in the concentration of CO₂(g) between time \( t_0 \) and time \( t_1 \). [1]
30. Base your answer on the information below.
The accompanying graph shows the relationship between the solubility of a sequence of primary alcohols in water and the total number of carbon atoms in a molecule of the corresponding alcohol at the same temperature and pressure. A primary alcohol has the --OH group located on an end carbon of the hydrocarbon chain.

a) Describe the relationship between the solubility of a primary alcohol in water and the total number of carbon atoms in the primary alcohol.
b) Determine the total mass of 1-pentanol that will dissolve in 110. grams of water to produce a saturated solution.

31. Base your answer on the accompanying graph. The graph shows the relationship between pH value and hydronium ion concentration for common aqueous solutions and mixtures. What is the hydronium ion concentration of tomato juice?
33. Base your answer on the information below:
The Balmer series refers to the visible bright lines in the spectrum produced by hydrogen atoms. The color and wavelength of each line in this series are given in the accompanying table.
On the diagram, draw *four* vertical lines to represent the Balmer series. [1]

**BALMER SERIES FOR HYDROGEN**

<table>
<thead>
<tr>
<th>Color</th>
<th>Wavelength (nm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>red</td>
<td>656.3</td>
</tr>
<tr>
<td>blue green</td>
<td>486.1</td>
</tr>
<tr>
<td>blue</td>
<td>434.1</td>
</tr>
<tr>
<td>violet</td>
<td>410.2</td>
</tr>
</tbody>
</table>

34. Base your answer on the information in the accompanying table. Based on the data in the table, state the relationship between the boiling point at 1 atmosphere and molar mass for these four substances. [1]

![Graph showing wavelength in nm vs. molar mass](image)

**Molar Mass and Boiling Point of Four Substances**

<table>
<thead>
<tr>
<th>Substance</th>
<th>Molar Mass (g/mol)</th>
<th>Boiling Point at 1 atm (K)</th>
</tr>
</thead>
<tbody>
<tr>
<td>methane</td>
<td>16</td>
<td>112</td>
</tr>
<tr>
<td>ethane</td>
<td>30</td>
<td>185</td>
</tr>
<tr>
<td>propane</td>
<td>44</td>
<td>231</td>
</tr>
<tr>
<td>butane</td>
<td>58</td>
<td>273</td>
</tr>
</tbody>
</table>

35. Base your answers on the information in the accompanying table.
   
   *On the grid shown or on a separate piece of paper, mark an appropriate scale on the axis labeled "Boiling Point (K)."* [1]

   *On the same grid, plot the data from the data table. Circle and connect the points.* [1]
36. Base your answer on the information below and the accompanying table. A substance is a solid at 15°C. A student heated a sample of the solid substance and recorded the temperature at one-minute intervals in the accompanying data table.

a. Copy the accompanying grid on a separate piece of paper, mark an appropriate scale on the axis labeled "Temperature (°C)." An appropriate scale is one that allows a trend to be seen.

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

c. Based on the data table, what is the melting point of this substance?
37. A student constructs a model for comparing the masses of subatomic particles. The student selects a small, metal sphere with a mass of 1 gram to represent an electron. A sphere with which mass would be most appropriate to represent a proton?

1. 1 g
2. 1/2 g
3. 1/2000 g
4. 2000 g

38. A plan is being developed for an experiment to test the effect of concentrated strong acids on a metal surface protected by various coatings. Some safety precautions would be the wearing of chemical safety goggles, an apron, and gloves. State one additional safety precaution that should be included in the plan. [1]