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Grades 9-12

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CHEMISTRY

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Name________________________________

Laboratory Dos And Don'ts

Identify what is wrong in each laboratory activity.

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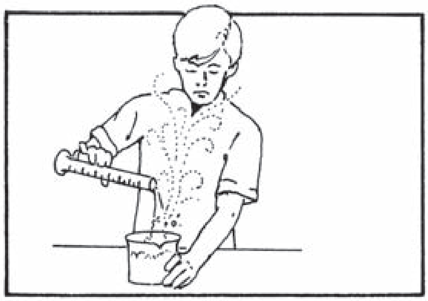

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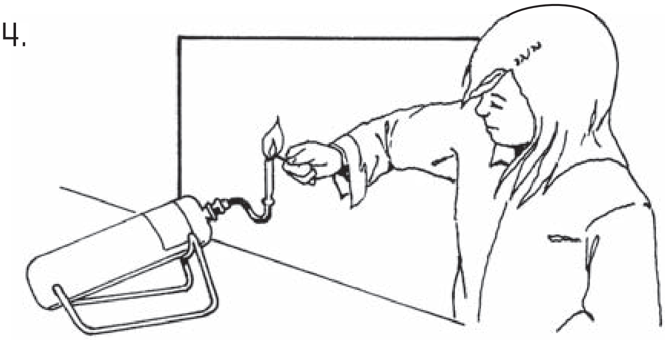

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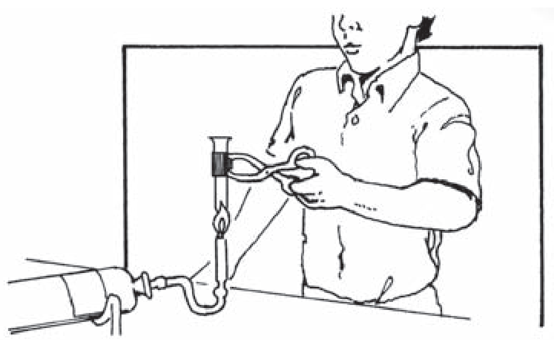

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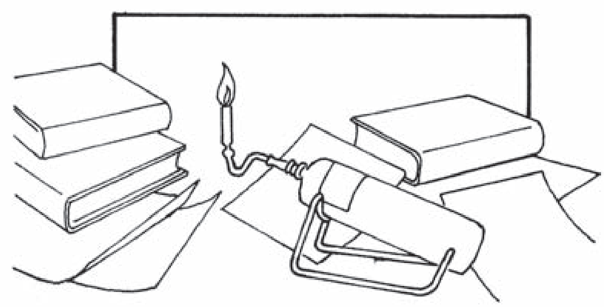

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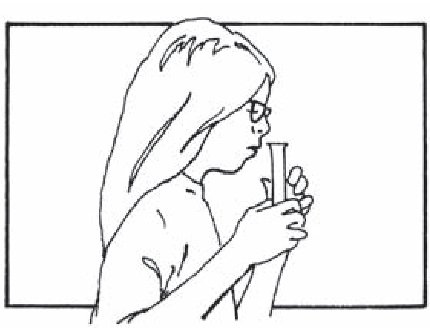

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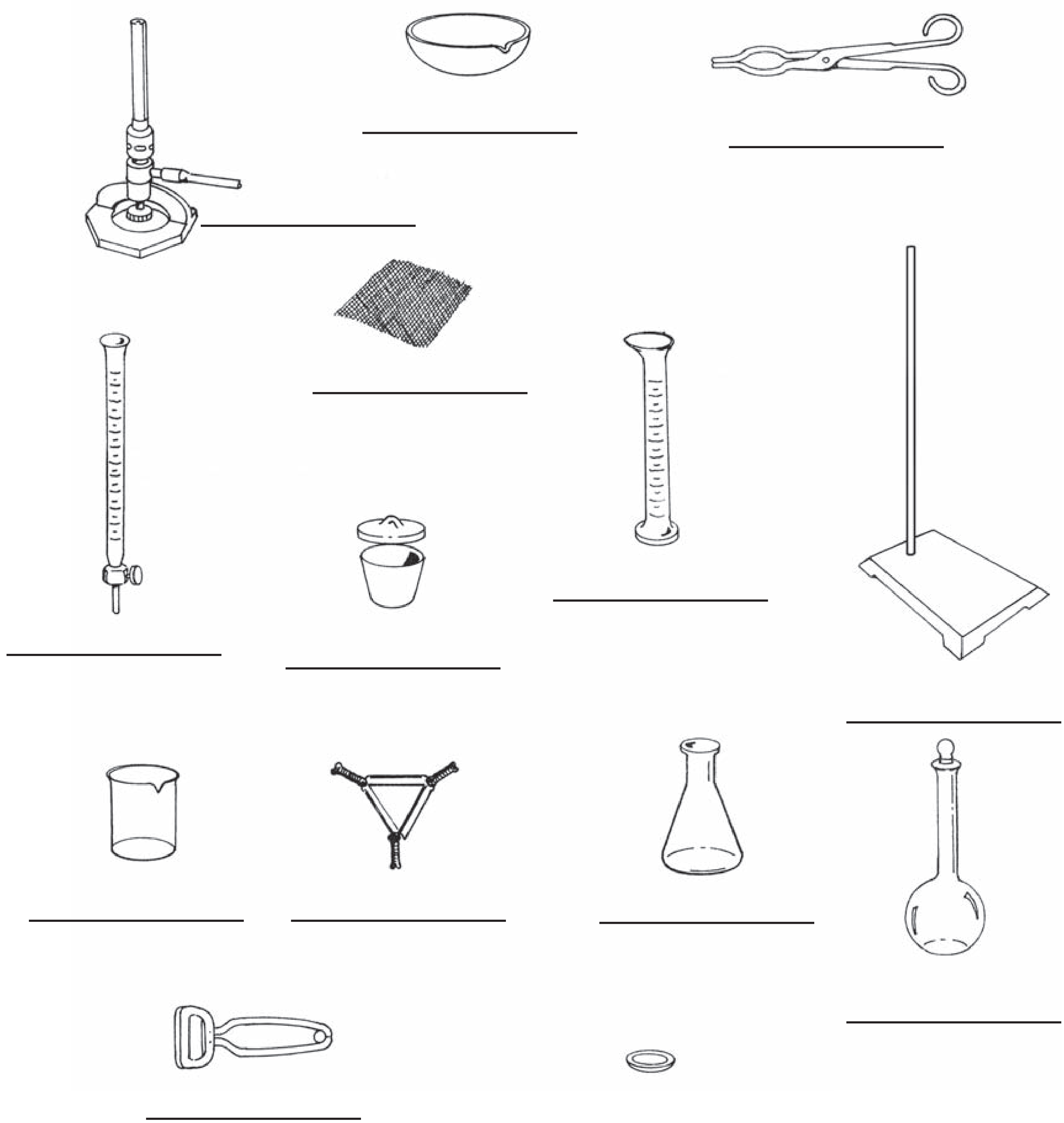
Name________________________________

Laboratory Equipment

Label the lab equipment.



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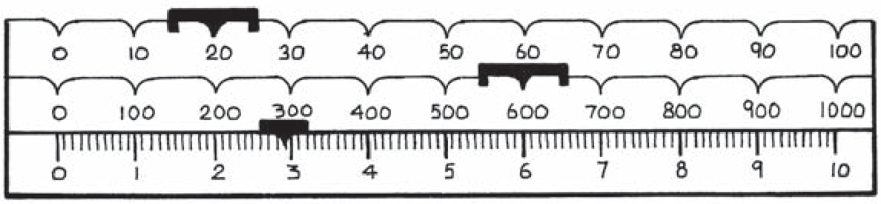
Triple And Four Beam Balances

Identify the mass on each balance.

Triple Beam Balance



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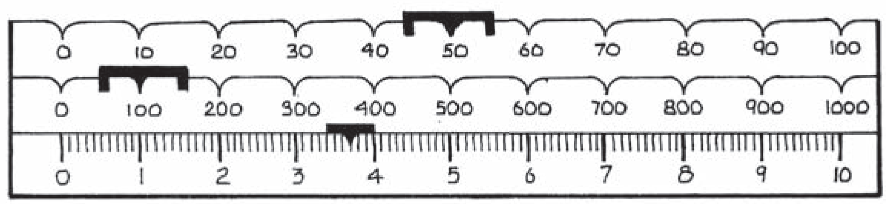

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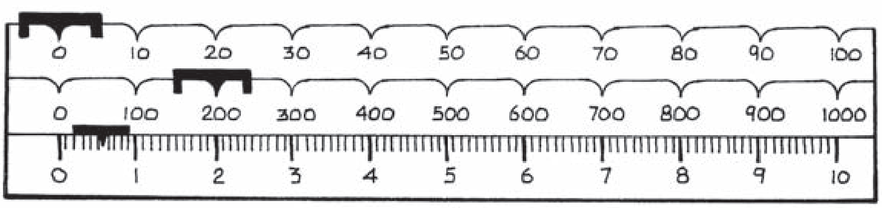

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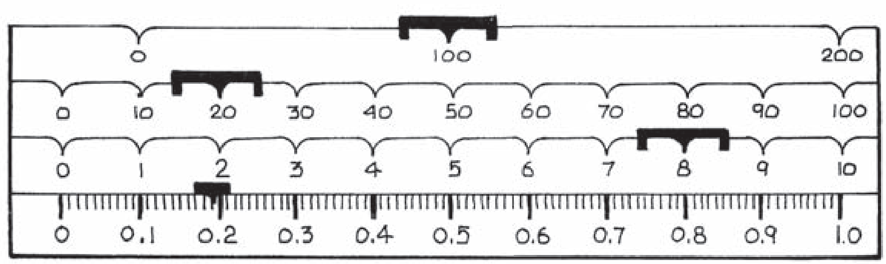
Four Beam Balance



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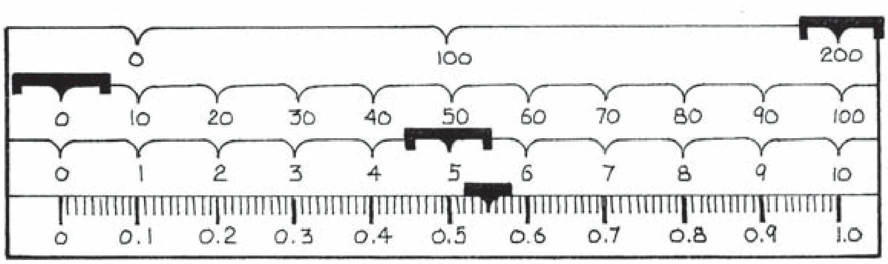

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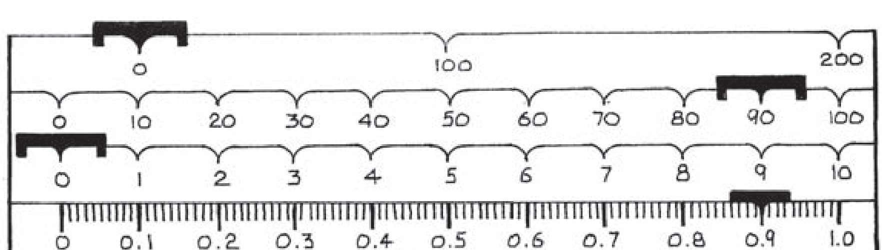

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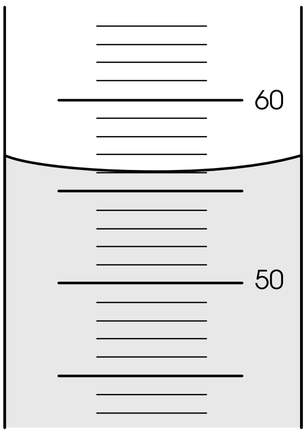
Measuring Liquid Volume

Identify the volume indicated on each graduated cylinder. The unit of volume is mL.



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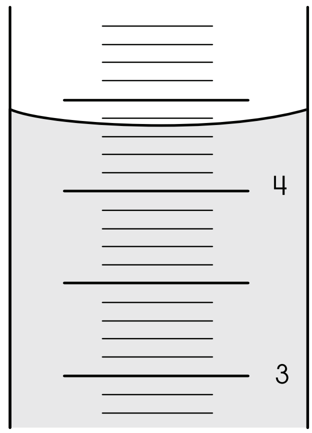

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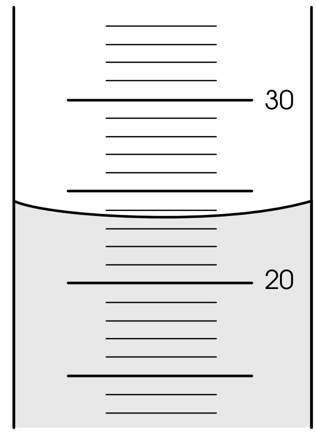

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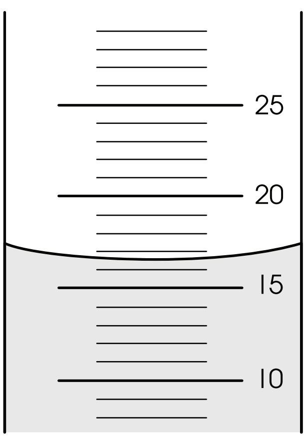

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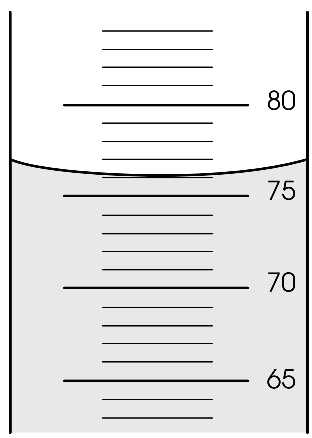

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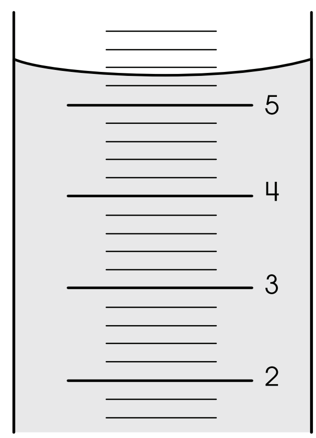

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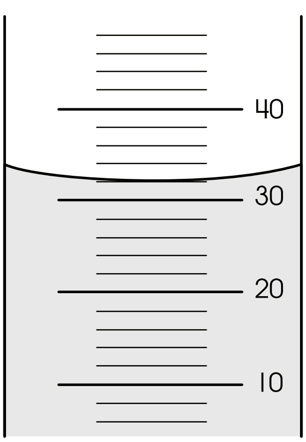

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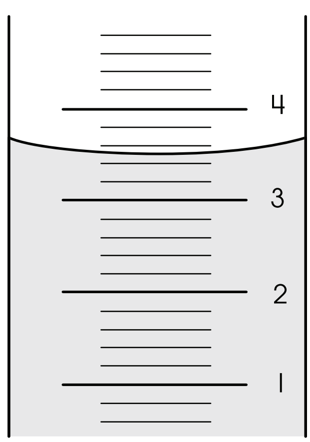

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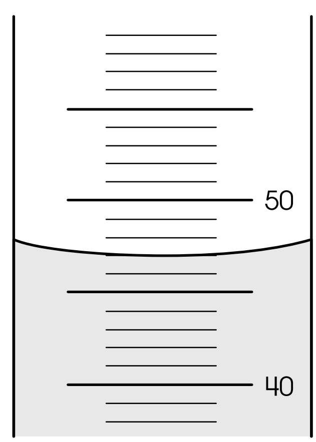

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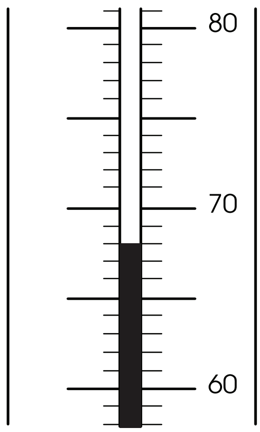
Name________________________________

Reading Thermometers

Identify the temperature indicated on each thermometer.



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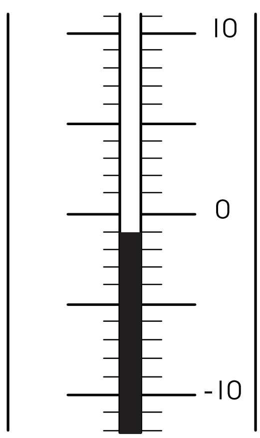

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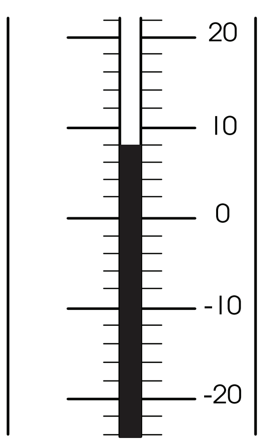

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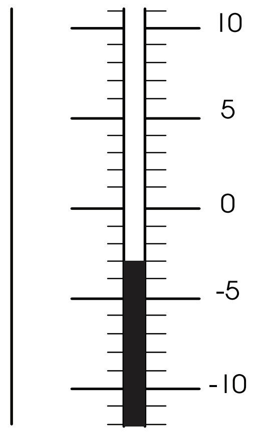

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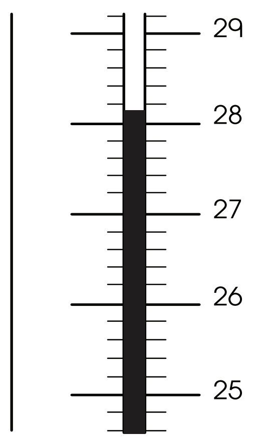

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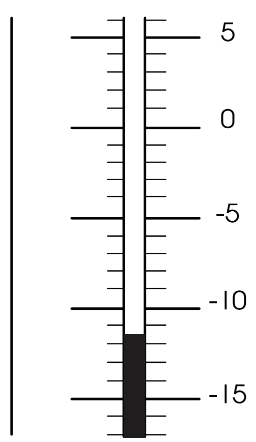

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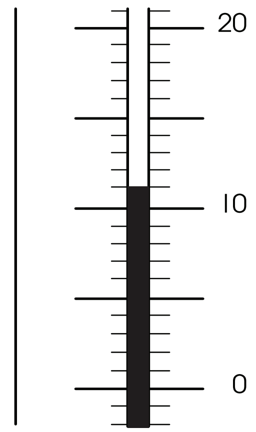

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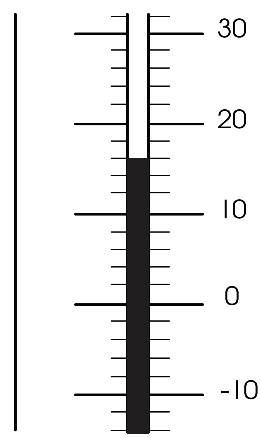

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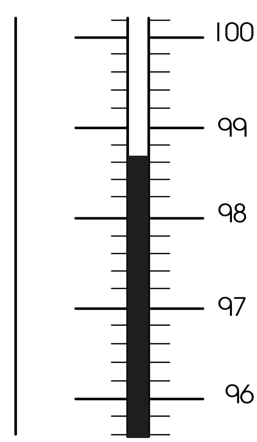


9. ____________________

Name________________________________

Metrics And Measurements

In the chemistry classroom and lab, the metric system of measurement is used. It is important to be able to convert from one unit to another.

mega-	kilo-	hecto-	deca-	Basic Units	deci-	centi-	milli-	micro-
(M)	(k)	(h)	(da)	gram (g)	(d)	(c)	(m)	(μ)
1,000,000	1,000	100	10	liter (L)	0.1	0.01	0.001	0.000001
10^6	10^3	10^2	10^1	meter (m)	10^{-1}	10^{-2}	10^{-3}	10^{-6}

Unit Factor Method

1. Write the given number and unit.
2. Set up a conversion factor (fraction used to convert one unit to another).
   a. Place the given unit as the denominator of the conversion factor.
   b. Place the desired unit as the numerator.
   c. Place a one in front of the larger unit.
   d. Determine the number of smaller units needed to make one of the larger units.
3. Cancel the units. Solve the problem.

Example 1: 55 mm = __________ m

$$55\ \cancel{mm} \times \frac{1\ m}{1{,}000\ \cancel{mm}} = 0.055\ m$$

Example 2: 88 km = __________ m

$$88\ \cancel{km} \times \frac{1{,}000\ m}{1\ \cancel{km}} = 88{,}000\ m$$

Example 3: 7,000 cm = __________ hm

$$7{,}000\ \cancel{cm} \times \frac{1\ \cancel{m}}{100\ \cancel{cm}} \times \frac{1\ hm}{10{,}000\ \cancel{cm}} = 0.7\ hm$$

Example 4: 8 daL = __________ dL

$$8\ \cancel{daL} \times \frac{10\ \cancel{L}}{1\ \cancel{daL}} \times \frac{100\ dL}{1\ \cancel{daL}} = 800\ dL$$

The unit factor method can be used to solve virtually any problem involving changes in units. It is especially useful in making complex conversions dealing with concentrations and derived units.

Convert each measurement.

1. 35 mL = __________ dL
2. 275 mm = __________ cm
3. 1,000 mL = __________ L
4. 25 cm = __________ mm
5. 0.075 m = __________ cm
6. 950 g = __________ kg
7. 1,000 L = __________ kL
8. 4,500 mg = __________ g
9. 0.005 kg = __________ dag
10. 15 g = __________ mg

Name________________________________

Dimensional Analysis (Unit Factor Method)

Using this method, it is possible to solve many problems by using the relationship of one unit to another. For example, 12 inches = one foot. Since these two numbers represent the same value, the fractions 12 in./1 ft. and 1 ft./12 in. are both equal to one. When you multiply another number by the number one, you do not change its value. However, you may change its unit.

Example 1: Convert 2 miles to inches.

$$2 \text{ miles} \times \frac{5{,}280 \text{ ft.}}{1 \text{ mile}} \times \frac{12 \text{ inches}}{1 \text{ ft.}} = 126{,}720 \text{ in.}$$

Example 2: How many seconds are in 4 days?

$$4 \text{ days} \times \frac{24 \text{ hrs.}}{1 \text{ day}} \times \frac{60 \text{ min.}}{1 \text{ hr.}} \times \frac{60 \text{ sec.}}{1 \text{ min.}} = 345{,}600 \text{ sec.}$$

Solve each problem. Round irrational numbers to the thousandths place.

1. 3 hr. = __________________ sec.
2. 0.035 mg = __________________ cg
3. 5.5 kg = __________________ lb.
4. 2.5 yd. = __________________ in.
5. 1.3 yr. = __________________ hr.
6. 3 moles = __________________ molecules (1 mole = 6.02×10^{23} molecules)
7. 2.5×10^{24} molecules = __________________ moles
8. 5 moles = __________________ liters (1 mole = 22.4 liters)
9. 100. liters = __________________ moles
10. 50. liters = __________________ molecules
11. 5.0×10^{24} molecules = __________________ liters
12. 7.5×10^{3} mL = __________________ liters

Name________________________________

Scientific Notation

Scientists very often deal with very small and very large numbers, which can lead to a lot of confusion when counting zeros. We can express these numbers as powers of 10.

Scientific notation takes the form of $M \times 10^n$ where $1 \leq M < 10$ and n represents the number of decimal places to be moved. Positive n indicates the standard form is a large number. Negative n indicates a number between zero and one.

Example 1: Convert 1,500,000 to scientific notation.

Move the decimal point so that there is only one digit to its left, for a total of 6 places.

$$1{,}500{,}000 = 1.5 \times 10^6$$

Example 2: Convert 0.000025 to scientific notation.

For this, move the decimal point 5 places to the right.

$$0.000025 = 2.5 \times 10^{-5}$$

(Note that when a number starts out less than one, the exponent is always negative.)

Convert each number to scientific notation.

1. 0.005 = ________________
2. 5,050 = ________________
3. 0.0008 = ________________
4. 1,000 = ________________
5. 1,000,000 = ________________
6. 0.25 = ________________
7. 0.025 = ________________
8. 0.0025 = ________________
9. 500 = ________________
10. 5,000 = ________________

Convert each number to standard notation.

11. 1.5×10^3 = ________________
12. 1.5×10^{-3} = ________________
13. 3.75×10^{-2} = ________________
14. 3.75×10^2 = ________________
15. 2.2×10^5 = ________________
16. 3.35×10^{-1} = ________________
17. 1.2×10^{-4} = ________________
18. 1×10^4 = ________________
19. 1×10^{-1} = ________________
20. 4×10^0 = ________________

Name______________________________

Significant Figures

A measurement can only be as accurate and precise as the instrument that produced it. A scientist must be able to express the accuracy of a number, not just its numerical value. We can determine the accuracy of a number by the number of significant figures it contains.

1. All digits 1–9 inclusive are significant.

   Example: 129 has 3 significant figures.

2. Zeros between significant digits are always significant.

   Example: 5,007 has 4 significant figures.

3. Trailing zeros in a number are significant only if the number contains a decimal point. Sometimes, a decimal may be added without any number in the tenths place.

   Example: 100.0 has 4 significant figures.
   100. has 3 significant figures.
   100 has 1 significant figure.

4. Zeros in the beginning of a number whose only function is to place the decimal point are not significant.

   Example: 0.0025 has 2 significant figures.

5. Zeros following a decimal significant figure are significant.

   Example: 0.000470 has 3 significant figures.
   0.47000 has 5 significant figures.

Determine the number of significant figures in each number.

1. 0.02 _________
2. 0.020 _________
3. 501 _________
4. 501.0 _________
5. 5,000 _________
6. 5,000. _________
7. 6,051.00 _________
8. 0.0005 _________
9. 0.1020 _________
10. 10,001 _________

Determine the location of the last significant place value by placing a bar over the digit. (Example: $1.70\overline{0}$)

11. 8,040
12. 0.90100
13. 3.01×10^{21}
14. 0.0300
15. 90,100
16. 0.000410
17. 699.5
18. 4.7×10^{-8}
19. 2.000×10^{2}
20. 10,800,000.0

Name________________________________

Calculations Using Significant Figures

When multiplying and dividing, limit and round to the least number of significant figures in any of the factors.

Example 1: $23.0\text{ cm} \times 432\text{ cm} \times 19\text{ cm} = 188{,}784\text{ cm}^3$

The answer is expressed as $190{,}000\text{ cm}^3$ since 19 cm has only two significant figures.

When adding and subtracting, limit and round your answer to the least number of decimal places in any of the numbers that make up your answer.

Example 2: 123.25 mL + 46.0 mL + 86.257 mL = 255.507 mL

The answer is expressed as 255.5 mL since 46.0 mL has only one decimal place.

Perform each operation, expressing the answer in the correct number of significant figures.

1. $1.35\text{ m} \times 2.467\text{ m} =$ ____________________
2. $1{,}035\text{ m}^2 \div 42\text{ m} =$ ____________________
3. 12.01 mL + 35.2 mL + 6 mL = ____________________
4. 55.46 g – 28.9 g = ____________________
5. $0.021\text{ cm} \times 3.2\text{ cm} \times 100.1\text{ cm} =$ ____________________
6. 0.15 cm + 1.15 cm + 2.051 cm = ____________________
7. $150\text{ L}^3 \div 4\text{ L} =$ ____________________
8. 505 kg – 450.25 kg = ____________________
9. $1.252\text{ mm} \times 0.115\text{ mm} \times 0.012\text{ mm} =$ ____________________
10. $1.278 \times 10^3\text{ m}^2 \div 1.4267 \times 10^2\text{ m} =$ ____________________

Name______________________________

Percentage Error

Percentage error is a way for scientists to express how far off a laboratory value is from the commonly accepted value.

The formula is:

$$\% \text{ error} = \left| \frac{\text{Accepted Value} - \text{Experimental Value}}{\text{Accepted Value}} \right| \times 100$$

absolute value →

Determine the percentage error in each problem.

1.	Experimental value = 1.24 g Accepted value = 1.30 g
2.	Experimental value = 1.24×10^{-2} g Accepted value = 9.98×10^{-3} g
3.	Experimental value = 252 mL Accepted value = 225 mL
4.	Experimental value = 22.2 L Accepted value = 22.4 L
5.	Experimental value = 125.2 mg Accepted value = 124.8 mg

Name____________________

Temperature and Its Measurement

Temperature (which measures average kinetic energy of the molecules) can be measured using three common scales: **Celsius**, **Kelvin**, and **Fahrenheit**. Use the following formulas to convert from one scale to another. Celsius is the scale most desirable for laboratory work. Kelvin represents the absolute scale. Fahrenheit is the old English scale, which is rarely used in laboratories.

$°C = K - 273$ $\quad$ $K = °C + 273$

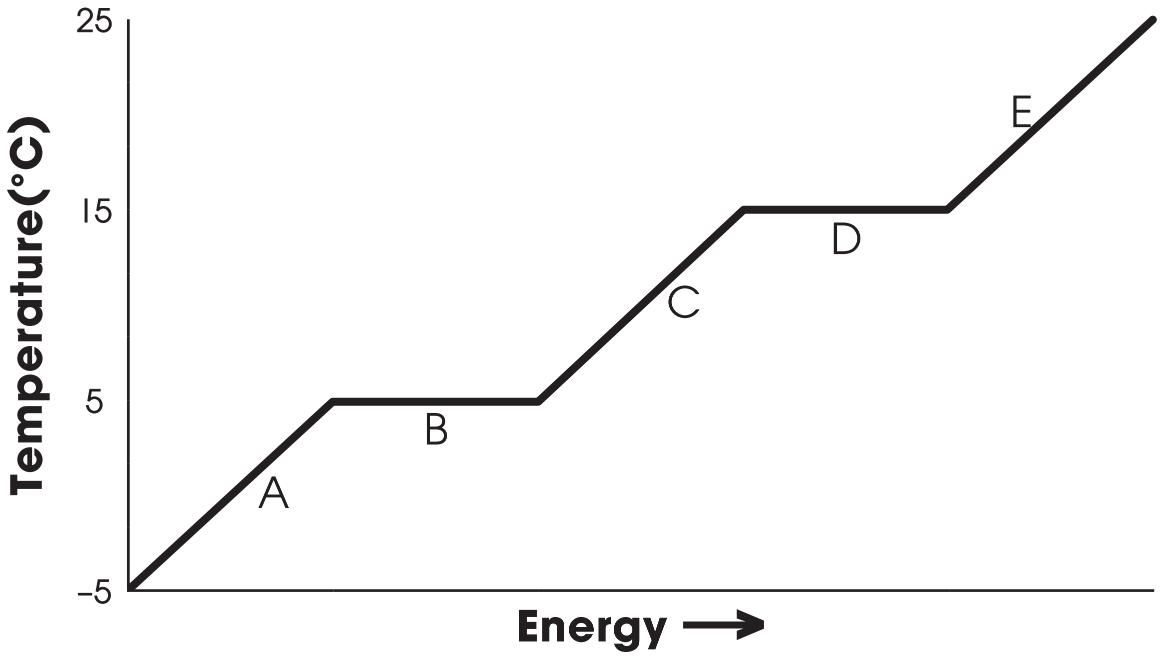
$°F = \frac{9}{5}\,°C + 32$ $\quad$ $°C = \frac{5}{9}(°F - 32)$

Complete the chart. All measurements are good to 1°C or better.

	°C	K	°F
1.	0°C		
2.			212°F
3.		450 K	
4.			98.6°F
5.	-273°C		
6.		294 K	
7.			77°F
8.		225 K	
9.	-40°C		

Name________________________________

Freezing and Boiling Point Graph

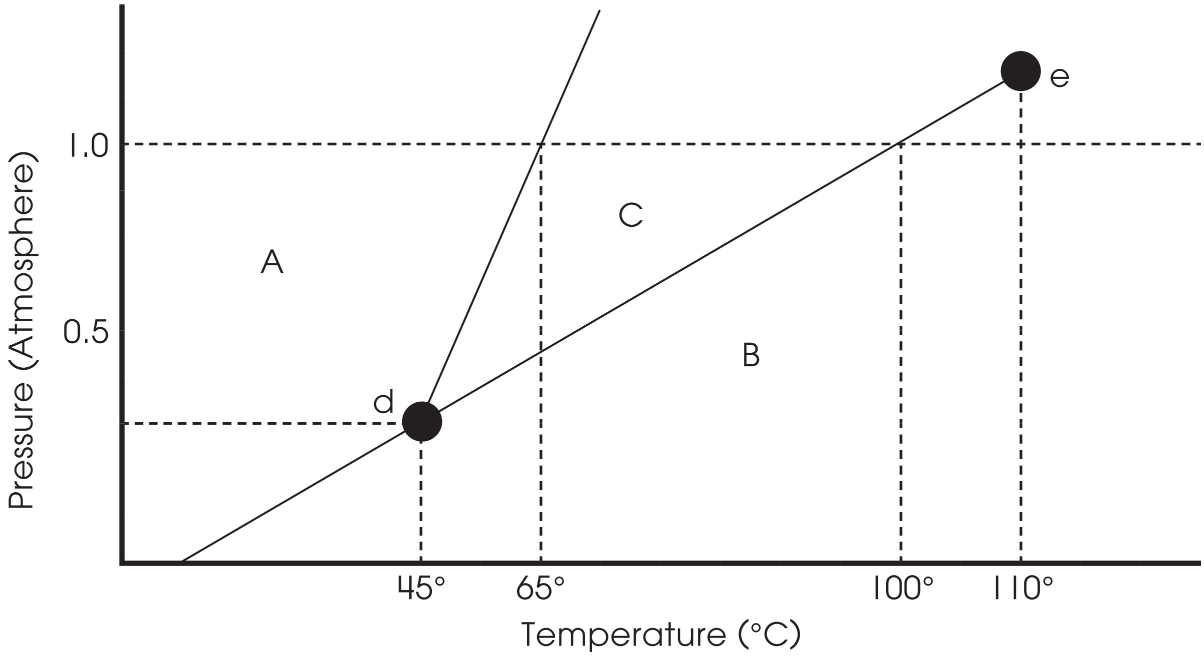


Use the graph to answer each question.

1. Which is the freezing point of the substance? ___________
2. Which is the boiling point of the substance? ___________
3. Which is the melting point of the substance? ___________
4. Which letter represents the range where the solid is being warmed? ___________
5. Which letter represents the range where the liquid is being warmed? ___________
6. Which letter represents the range where the vapor is being warmed? ___________
7. Which letter represents the melting of the solid? ___________
8. Which letter represents the vaporization of the liquid? ___________
9. Which letter(s) shows a change in potential energy? ___________
10. Which letter(s) shows a change in kinetic energy? ___________
11. Which letter represents condensation? ___________
12. Which letter represents crystallization? ___________

Name________________________________

Phase Diagram



Use the diagram to answer each question.

1. Which section represents the solid phase? __________
2. Which section represents the liquid phase? __________
3. Which section represents the gas phase? __________
4. Which letter represents the triple point? __________
5. Which letter represents the critical point? __________
6. What is this substance's normal melting point? __________
7. What is this substance's normal boiling point? __________
8. Above what temperature is it impossible to liquify this substance no matter what the pressure? __________
9. At what temperature and pressure do all three phases coexist? __________________
10. Is the density of the solid greater than or less than the density of the liquid?

    ____

11. Would an increase in pressure cause this substance to freeze or melt? ____________

Name____________________________

Heat and Its Measurement

Heat (or energy) can be measured in units of **calories** or **joules**. When there is a temperature change (ΔT), heat (Q) can be calculated using this formula:

Q = mass × ΔT × specific heat capacity
(ΔT = final temperature – initial temperature)

During a phase change, use this formula:

Q = mass × heat of fusion (or heat of vaporization)

Solve each problem.

1. How many joules of heat are given off when 5.0 g of water cool from 75°C to 25°C? (Specific heat of water = 4.18 J/g°C)

2. How many calories are given off by the water in problem 1? (Specific heat of water = 1.0 cal/g°C)

3. How many joules does it take to melt 35 g of ice at O°C? (heat of fusion = 333 J/g)

4. How many calories are given off when 85 g of steam condense to liquid water? (heat of vaporization = 539.4 cal/g)

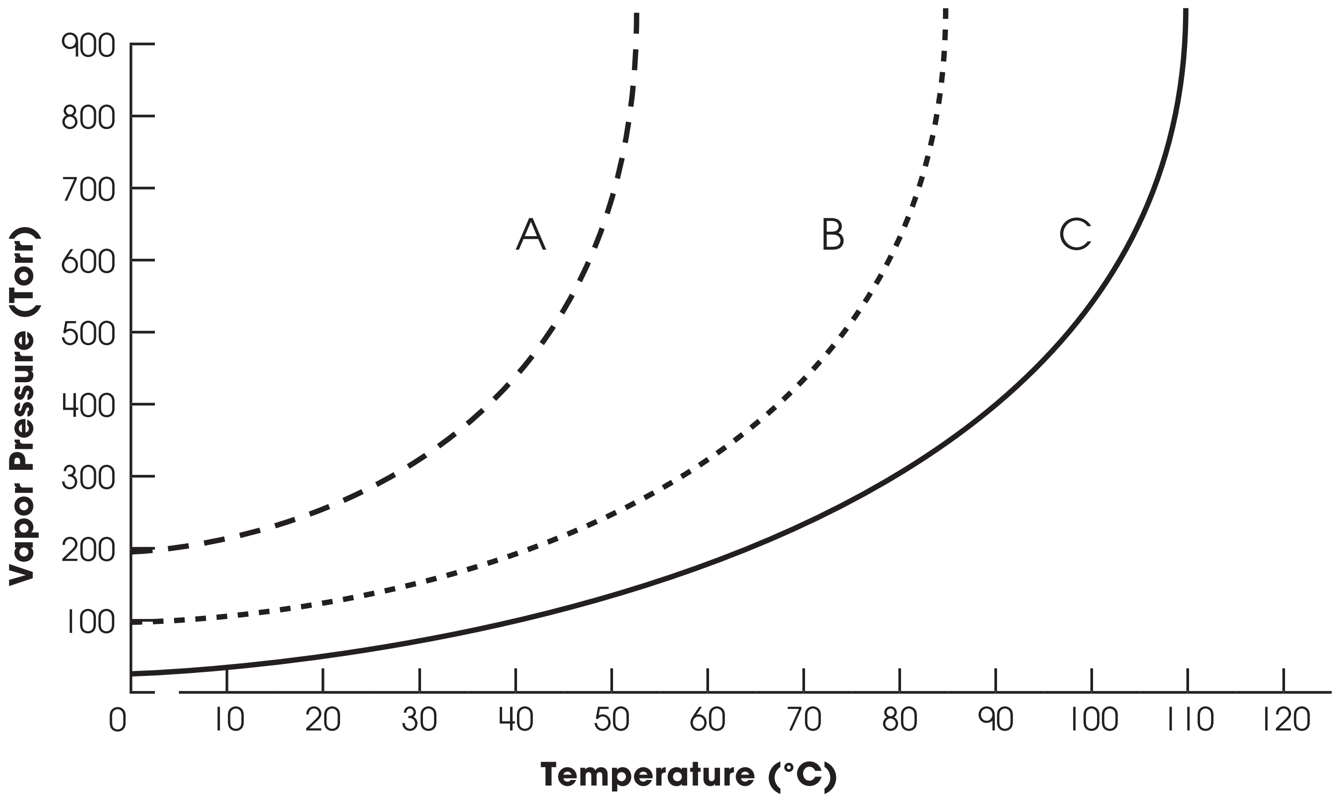
5. How many joules of heat are necessary to raise the temperature of 25 g of water from 10°C to 60°C?

6. How many calories are given off when 50 g of water at 0°C freezes? (heat of fusion = 79.72 cal/g)

Name______________________________

Vapor Pressure and Boiling

A liquid will boil when its vapor pressure equals the atmospheric pressure.



Use the graph to answer each question.

1. At what temperature would Liquid A boil at an atmospheric pressure of 400 Torr? _________
2. Liquid B? _________
3. Liquid C? _________
4. How low must the atmospheric pressure be for Liquid A to boil at 35°C? _________
5. Liquid B? _________
6. Liquid C? _________
7. What is the normal boiling point of Liquid A? _________
8. Liquid B? _________
9. Liquid C? _________
10. Which liquid has the strongest intermolecular forces? _________

Name________________________________

Matter—Substances vs. Mixtures

All matter can be classified as either a **substance** (element or compound) or a **mixture** (heterogeneous or homogeneous).

Matter

- **Substance**
  can write chemical formula, homogeneous
  - **Element**
    one type of atom
  - **Compound**
    two or more different atoms, chemically bonded
- **Mixture**
  variable ratio
  - **Homogeneous**
    solutions
  - **Heterogeneous**
    colloids and suspensions

Classify each of the following as a substance or a mixture. If it is a substance, write *element* or *compound* in the substance column. If it is a mixture, write *heterogeneous* or *homogeneous* in the mixture column.

Type of Matter	Substance	Mixture
1. chlorine		
2. water		
3. soil		
4. sugar water		
5. oxygen		
6. carbon dioxide		
7. rocky road ice cream		
8. alcohol		
9. pure air		
10. iron		

Name________________________________

Physical vs. Chemical Properties

A **physical property** is observed with the senses and can be determined without destroying the object. Color, shape, mass, length, and odor are all examples of physical properties.

A **chemical property** indicates how a substance reacts with something else. The original substance is fundamentally changed in observing a chemical property. For example, the ability of iron to rust is a chemical property. The iron has reacted with oxygen, and the original iron metal is changed. It now exists as iron oxide, a different substance.

Classify each property as either chemical or physical by putting a check in the appropriate column.

	Physical Property	Chemical Property
1. blue color		
2. density		
3. flammability		
4. solubility		
5. reacts with acid to form H_2		
6. supports combustion		
7. sour taste		
8. melting point		
9. reacts with water to form a gas		
10. reacts with a base to form water		
11. hardness		
12. boiling point		
13. can neutralize a base		
14. luster		
15. odor		

Name___________________________

Physical vs. Chemical Changes

In a **physical change**, the original substance still exists; it only changes in form. In a **chemical change**, a new substance is produced. Energy changes always accompany chemical changes.

Classify each as a *physical* or *chemical* change.

1. Sodium hydroxide dissolves in water. ____________________
2. Hydrochloric acid reacts with potassium hydroxide to produce a salt, water, and heat. ____________________
3. A pellet of sodium is sliced in two. ____________________
4. Water is heated and changed to steam. ____________________
5. Potassium chlorate decomposes to potassium chloride and oxygen gas. ____________________
6. Iron rusts. ____________________
7. When placed in H_2O, a sodium pellet catches on fire as hydrogen gas is liberated and sodium hydroxide forms. ____________________
8. Water evaporates. ____________________
9. Ice melts. ____________________
10. Milk sours. ____________________
11. Sugar dissolves in water. ____________________
12. Wood rots. ____________________
13. Pancakes are cooking on a griddle. ____________________
14. Grass is growing in a lawn. ____________________
15. A tire is inflated with air. ____________________
16. Food is digested in the stomach. ____________________
17. Water is absorbed by a paper towel. ____________________

Name________________________________

Boyle's Law

Boyle's Law states that the volume of a given sample of gas at a constant temperature varies inversely with the pressure. (If one goes up, the other goes down.) Use the formula:

$$P_1 \times V_1 = P_2 \times V_2$$

Solve each problem (assuming constant temperature).

1. A sample of oxygen gas occupies a volume of 250. mL at 740. Torr. What volume will it occupy at 800. Torr pressure?

   ____

2. A sample of carbon dioxide occupies a volume of 3.50 liters at 125 kPa pressure. What pressure would the gas exert if the volume was decreased to 2.00 liters?

   ____

3. A 2.0 liter container of nitrogen has a pressure of 3.2 atm. What volume would be necessary to decrease the pressure to 1.0 atm?

   ____

4. Ammonia gas occupies a volume of 450. mL at a pressure of 720. mmHg. What volume will it occupy at standard pressure?

   ____

5. A 175 mL sample of neon has its pressure changed from 75 kPa to 150 kPa. What is its new volume?

   ____

6. A sample of hydrogen at 1.5 atm has its pressure decreased to 0.50 atm, producing a new volume of 750 mL. What was its original volume?

   ____

7. Chlorine gas occupies a volume of 1.2 liters at 720 Torr. What volume will it occupy at 1 atm pressure?

   ____

8. Fluorine gas exerts a pressure of 900. Torr. When the pressure is changed to 1.50 atm, its volume is 250. mL. What was the original volume?

   ____

Name________________________________

Charles' Law

Charles' Law states that the volume of a given sample of gas at a constant pressure is directly proportional to the temperature in Kelvin. Use the following formulas:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{or} \quad V_1 \times T_2 = V_2 \times T_1$$

$$K = {}^\circ C + 273$$

Solve each problem (assuming constant pressure).

1. A sample of nitrogen occupies a volume of 250 mL at 25°C. What volume will it occupy at 95°C?

    ____

2. Oxygen gas is at a temperature of 40°C when it occupies a volume of 2.3 liters. To what temperature should it be raised to occupy a volume of 6.5 liters?

    ____

3. Hydrogen gas was cooled from $1\overline{5}0$°C to $5\overline{0}$°C. Its new volume is 75 mL. What was its original volume?

    ____

4. Chlorine gas occupies a volume of 25 mL at $3\overline{0}0$ K. What volume will it occupy at 600 K?

    ____

5. A sample of neon gas at $5\overline{0}$°C and a volume of 2.5 liters is cooled to 25°C. What is the new volume?

    ____

6. Fluorine gas at $30\overline{0}$ K occupies a volume of $50\overline{0}$ mL. To what temperature should it be lowered to bring the volume to $30\overline{0}$ mL?

    ____

7. Helium occupies a volume of 3.8 liters at -45°C. What volume will it occupy at 45°C?

    ____

8. A sample of argon gas is cooled and its volume went from $38\overline{0}$ mL to $25\overline{0}$ mL. If its final temperature was -55°C, what was its original temperature?

    ____

Name________________________________

Combined Gas Law

In practical terms, it is often difficult to hold any of the variables constant. When there is a change in pressure, volume, and temperature, the combined gas law is used.

$$\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2} \quad \text{or} \quad P_1 V_1 T_2 = P_2 V_2 T_1$$

Complete the chart.

	P_1	V_1	T_1	P_2	V_2	T_2
1.	1.5 atm	3.0 L	2$\bar{0}$°C	2.5 atm		3$\bar{0}$°C
2.	720 Torr	256 mL	25°C		250 mL	5$\bar{0}$°C
3.	6$\bar{0}$0 mmHg	2.5 L	22°C	760 mmHg	1.8 L	
4.		750 mL	0.0°C	2.0 atm	500 mL	25°C
5.	95 kPa	4.0 L		101 kPa	6.0 L	471 K or 198°C
6.	650. Torr		100°	900. Torr	225 mL	15$\bar{0}$°C
7.	850 mmHg	1.5 L	15°C		2.5 L	3$\bar{0}$°C
8.	125 kPa	125 mL		10$\bar{0}$ kPa	10$\bar{0}$ mL	75°C

Name________________________________

Dalton's Law of Partial Pressures

Dalton's Law says that the sum of the individual pressures of all the gases that make up a mixture is equal to the total pressure, or: $P_T = P_1 + P_2 + P_3 + \ldots$ The partial pressure of each gas is equal to the mole fraction of each gas times the total pressure.

$$P_T = P_1 + P_2 + P_3 + \ldots \quad \text{or} \quad \frac{\text{moles gas}_x}{\text{total moles}} \times P_T = P_X$$

Solve each problem.

1. A 250. mL sample of oxygen is collected over water at 25°C and 760.0 Torr. What is the pressure of the dry gas alone? (Vapor pressure of water at 25°C = 23.8 Torr)

2. A 32.0 mL sample of hydrogen is collected over water at $2\overline{0}$°C and 750.0 Torr. What is the pressure of the dry gas alone? (Vapor pressure of water at $2\overline{0}$°C = 17.5 Torr)

3. A 54.0 mL sample of oxygen is collected over water at 23°C and 770.0 Torr. What is the pressure of the dry gas alone? (Vapor pressure of water at 23°C = 21.1 Torr)

4. A mixture of 2.00 moles of H_2, 3.00 moles of NH_3, 4.00 moles of CO_2, and 5.00 moles of N_2 exerts a total pressure of $80\overline{0}$ Torr. What is the partial pressure of each gas?

5. The partial pressure of F_2 is 300 Torr in a mixture of gases where the total pressure is 1.00 atm. If there are 1.5 total moles in the mixture, how many moles of F_2 are present?

Name ______________________________

Ideal Gas Law

The **ideal gas law** describes the state of an ideal gas. While an ideal gas is hypothetical, the ideal gas law can be used to approximate the behavior of many gases under normal conditions. Use the formula:

$PV = nRT$ where

P = pressure in atmospheres

V = volume in liters

R = Universal Gas Constant 0.0821 L•atm/mol•K

n = number of moles of gas

T = Kelvin temperature

Use the ideal gas law to solve each problem.

1. How many moles of oxygen will occupy a volume of 2.5 liters at 1.2 atm and 25°C?

   ____

2. What volume will 2.0 moles of nitrogen occupy at 720 Torr and $2\bar{0}$°C?

   ____

3. What pressure will be exerted by 25 g of CO_2 at a temperature of 25°C and a volume of $50\bar{0}$ mL? ____________________

4. At what temperature will 5.00 g of Cl_2 exert a pressure of 900. Torr at a volume of $75\bar{0}$ mL? ____________________

5. What is the density of NH_3 at $80\bar{0}$ Torr and 25°C? ____________________

6. If the density of a gas is 1.2 g/L at 745. Torr and $2\bar{0}$°C, what is its molecular mass?

   ____

7. How many moles of nitrogen gas will occupy a volume of 347 mL at 6680 Torr and 27°C? ____________________

8. What volume will 454 grams (1 lb.) of hydrogen occupy at 1.05 atm and 25°C?

   ____

9. Find the number of grams of CO_2 that exert a pressure of 785 Torr at a volume of 32.5 L and a temperature of 32°C. ____________________

10. An elemental gas has a mass of 10.3 g. If the volume is 58.4 L and the pressure is 758 Torr at a temperature of 2.5°C, what is the gas? ____________________

Name______________________________

Graham's Law of Effusion

Graham's Law states that a gas will effuse at a rate that is inversely proportional to the square root of its molecular mass, *MM*.

$$\frac{rate_1}{rate_2} = \sqrt{\frac{MM_2}{MM_1}}$$

Solve each problem.

1.	Under the same conditions of temperature and pressure, how many times faster will hydrogen effuse compared to carbon dioxide?
2.	If the carbon dioxide in problem 1 takes 32 seconds to effuse, how long will the hydrogen take?
3.	What is the relative rate of effusion of NH_3 compared to helium? Does NH_3 effuse faster or slower than helium?
4.	If the helium in problem 3 takes 20 seconds to effuse, how long will NH_3 take?
5.	An unknown gas effuses 0.25 times as fast as helium. What is the molecular mass of the unkown gas?

Name________________________________

Element Symbols

An element symbol can stand for one atom of the element or one mole of atoms of the element. (One mole = 6.02×10^{23} atoms of an element.)

Write the symbol for each element.

1. oxygen __________
2. hydrogen __________
3. chlorine __________
4. mercury __________
5. fluorine __________
6. barium __________
7. helium __________
8. uranium __________
9. radon __________
10. sulfur __________
11. plutonium __________
12. calcium __________
13. radium __________
14. cobalt __________
15. zinc __________
16. arsenic __________
17. lead __________
18. iron __________

Write the name of the element that corresponds with each symbol.

19. Kr __________________________
20. K __________________________
21. C __________________________
22. Ne __________________________
23. Si __________________________
24. Zr __________________________
25. Sn __________________________
26. Pt __________________________
27. Na __________________________
28. Al __________________________
29. Cu __________________________
30. Ag __________________________
31. P __________________________
32. Mn __________________________
33. I __________________________
34. Au __________________________
35. Mg __________________________
36. Ni __________________________
37. Br __________________________
38. Hg __________________________

Name____________________________

Atomic Structure

An atom is made up of protons and neutrons (both found in the nucleus) and electrons (in the surrounding electron cloud). The **atomic number** is equal to the number of protons. The **mass number** is equal to the number of protons plus neutrons.

In a neutral atom, the number of protons equals the number of electrons. The **charge** on an ion indicates an imbalance between protons and electrons. Too many electrons produce a negative charge. Too few electrons produce a positive charge.

This structure can be written as part of a chemical symbol.

mass number → $^{15}_{7}\mathrm{N}^{3+}$ ← *charge*

atomic number ↗

7 protons

8 neutrons (15 – 7)

4 electrons

Complete the chart.

Element/ Ion	Atomic Number	Atomic Mass	Mass Number	Protons	Neutrons	Electrons
H						
H^{+}						
$^{12}_{6}C$						
$^{7}_{3}Li^{+}$						
$^{35}_{17}Cl^{-}$						
$^{39}_{19}K$						
$^{24}_{12}Mg^{2+}$						
As^{3-}						
Ag						
Ag^{1+}						
S^{2-}						
U						

Name________________________________

Isotopes and Average Atomic Mass

Elements come in a variety of **isotopes**, meaning they are made up of atoms with the same atomic number but different atomic masses. These atoms differ in the number of neutrons.

The **average atomic mass** is the weighted average of all of the isotopes of an element.

Example: A sample of cesium is 75% ^{133}Cs, 20% ^{132}Cs, and 5% ^{134}Cs. What is its average atomic mass?

Answer: $0.75 \times 133 = 99.75$
$0.20 \times 132 = 26.4$
$0.05 \times 134 = \underline{6.7}$
Total = 132.85 amu = average atomic mass

Determine the average atomic mass of each mixture of isotopes.

1.	80% ^{127}I, 17% ^{126}I, 3% ^{128}I
2.	50% ^{197}Au, 50% ^{198}Au
3.	15% ^{55}Fe, 85% ^{56}Fe
4.	99% ^{1}H, 0.8% ^{2}H, 0.2% ^{3}H
5.	95% ^{14}N, 3% ^{15}N, 2% ^{16}N
6.	98% ^{12}C, 2% ^{14}C

Name________________________________

Electron Configuration (Level One)

Electrons are distributed in the electron cloud into principal energy levels (1, 2, 3, ...), sublevels (*s, p, d, f*), orbitals (*s* has 1, *p* has 3, *d* has 5, *f* has 7), and spin (two electrons allowed per orbital).

Example: Draw the electron configuration of sodium (atomic number 11).

$1s^2$	$2s^2$		$2p^6$		$3s^1$
↑↓	↑↓	↑↓	↑↓	↑↓	↑

Draw the electron configuration of each atom.

1. Cl

2. N

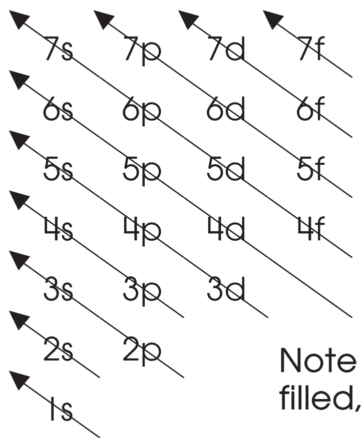
3. Al

4. O

Name________________________________

Electron Configuration (Level Two)

At atomic numbers greater than 18, the sublevels begin to fill out of order. A good approximation of the order of filling can be determined using the diagonal rule.



Note that after the 3p sublevel is filled, the 4s is filled, and then the 3d.

Draw the electron configuration of each atom.

1.	K
2.	V
3.	Co
4.	Zr

Name________________________________

Valence Electrons

The **valence electrons** are the electrons in the outermost principal energy level. They are always *s* electrons or *s* and *p* electrons. Since the total number of electrons possible in *s* and *p* sublevels is eight, there can be no more than eight valence electrons.

Example: carbon

Electron configuration is $1s^2$ $\boxed{2s^2 \quad 2p^2}$.

Carbon has 4 valence electrons.

Determine the number of valence electrons in each atom.

1. fluorine ____________
2. phosphorus ____________
3. calcium ____________
4. nitrogen ____________
5. iron ____________
6. argon ____________
7. potassium ____________
8. helium ____________
9. magnesium ____________
10. sulfur ____________
11. lithium ____________
12. zinc ____________
13. carbon ____________
14. iodine ____________
15. oxygen ____________
16. barium ____________
17. aluminum ____________
18. hydrogen ____________
19. xenon ____________
20. copper ____________

Name________________________________

Lewis Dot Diagrams

Lewis dot diagrams are a way to indicate the number of valence electrons around an atom.

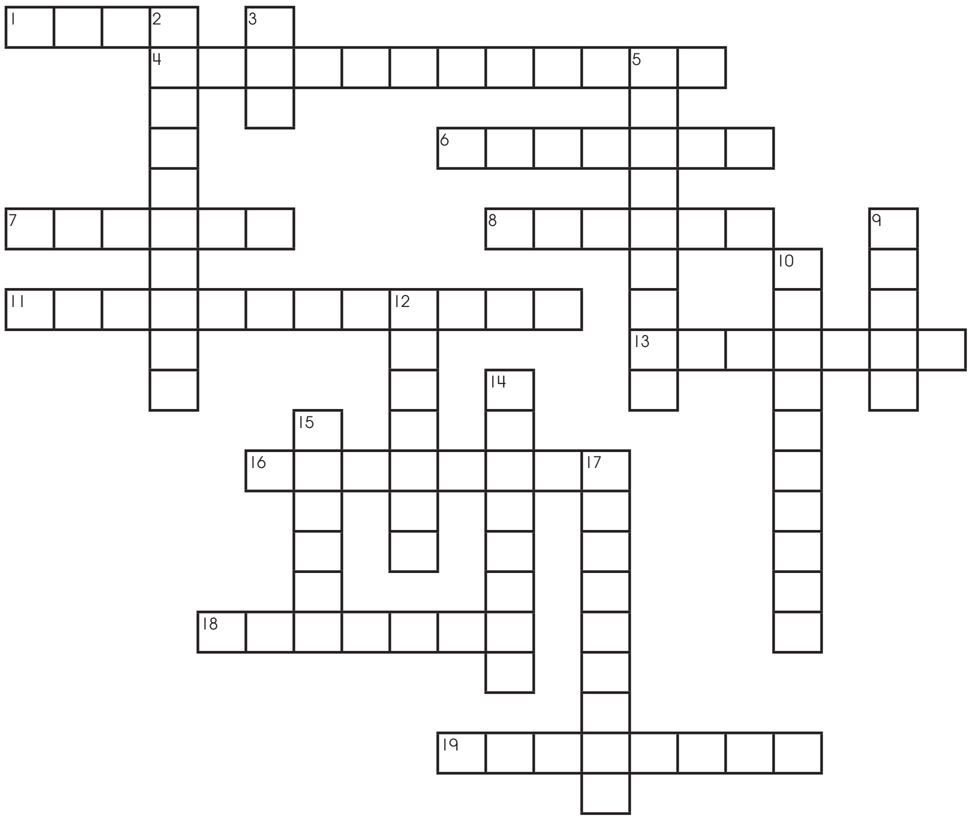
Examples:

Draw the Lewis dot diagram of each atom.

1. calcium
2. potassium
3. argon
4. aluminum
5. bromine
6. carbon
7. helium
8. oxygen
9. phosphorus
10. hydrogen

Name________________________________

Atomic Structure Crossword



Across

1. The smallest particle of an element that can enter into chemical change
4. The number of protons in the nucleus of an atom
6. Cannot be decomposed into simpler substances by ordinary chemical means
7. State in which all electrons are at their lowest possible energy level
8. The positively charged particle found in the nucleus
11. Standard atomic mass unit for carbon
13. Most of the mass of an atom is here.
16. Mass number minus atomic number
18. Electrons in the outermost principal energy level
19. Protons and neutrons are these.

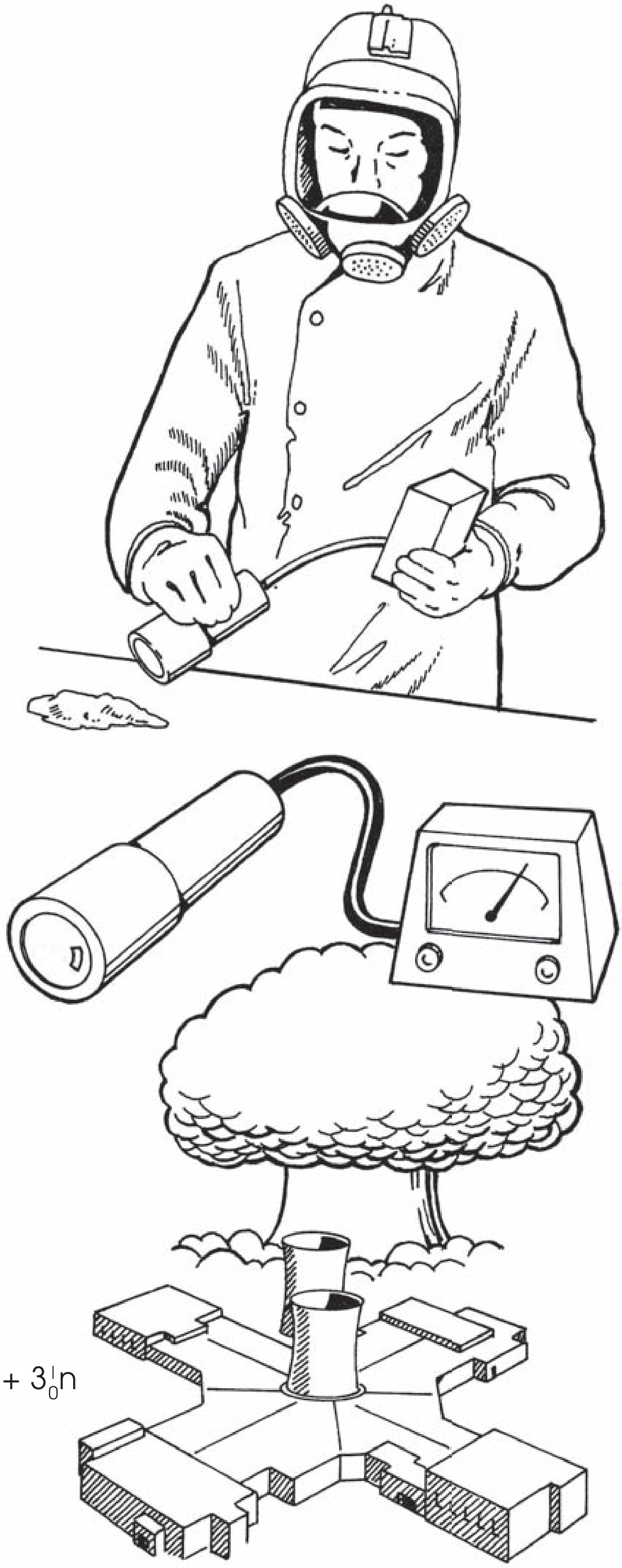
Down

2. Sum of the protons and neutrons in the nucleus of an atom
3. Charged atom or group of atoms
5. Equal to the number of protons in a neutral atom
9. The volume of an atom is determined by the size of its electron.
10. Different forms of the same element
12. State in which electrons have absorbed energy and "jumped" to a higher energy level
14. Atoms with the same atomic number but different atomic masses
15. The nucleus and all electrons in an atom except the valence electrons
17. *s, p, d, f*

Name________________________________

Nuclear Decay

Predict the product of each nuclear reaction.



1. $^{42}\mathrm{K} \longrightarrow {}^{0}_{-1}\mathrm{e} +$ ____________________

2. $^{239}\mathrm{Pu} \longrightarrow {}^{4}_{2}\mathrm{He} +$ ____________________

3. ${}^{235}_{92}\mathrm{U} \longrightarrow$ ____________________ $+ {}^{231}_{90}\mathrm{Th}$

4. ${}^{1}_{1}\mathrm{H} + {}^{3}_{1}\mathrm{H} \longrightarrow$ ____________________

5. ${}^{6}_{3}\mathrm{Li} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{4}_{2}\mathrm{He} +$ ____________________

6. ${}^{27}_{13}\mathrm{Al} + {}^{4}_{2}\mathrm{He} \longrightarrow {}^{30}_{15}\mathrm{P} +$ ____________________

7. ${}^{9}_{4}\mathrm{Be} + {}^{1}_{1}\mathrm{H} \longrightarrow$ ____________________ $+ {}^{4}_{2}\mathrm{He}$

8. $^{37}\mathrm{K} \longrightarrow {}^{0}_{+1}\mathrm{e} +$ ____________________

9. ____________________ $+ {}^{1}_{0}\mathrm{n} \longrightarrow {}^{142}_{56}\mathrm{Ba} + {}^{91}_{36}\mathrm{Kr} + 3{}^{1}_{0}\mathrm{n}$

10. ${}^{238}_{92}\mathrm{U} + {}^{4}_{2}\mathrm{He} \longrightarrow$ ____________________ $+ {}^{1}_{0}\mathrm{n}$

Name________________________________

Half-Lives of Radioactive Isotopes

Solve each problem.

1. How much of a 100.0 g sample of ^{198}Au is left after 8.10 days if its half-life is 2.70 days?

2. A 50.0 g sample of ^{16}N decays to 12.5 g in 14.4 seconds. What is its half-life?

3. The half-life of ^{42}K is 12.4 hours. How much of a $75\overline{0}$ g sample is left after 62.0 hours?

4. What is the half-life of ^{99}Tc if a $50\overline{0}$ g sample decays to 62.5 g in 639,000 years?

5. The half-life of ^{232}Th is 1.4×10^{10} years. If there are 25.0 g of the sample left after 2.8×10^{10} years, how many grams were in the original sample?

6. There are 5.0 g of ^{131}I left after 40.35 days. How many grams were in the original sample if its half-life is 8.07 days?

Name________________________________

Periodic Table Worksheet

Use a copy of the periodic table to answer each question.

1. Where are the most active metals located? ______________________________
2. Where are the most active nonmetals located? ___________________________
3. As you go from left to right across a period, the atomic size (decreases, increases).
   Why? ____
4. As you travel down a group, the atomic size (decreases, increases). Why?
   ____
5. A negative ion is (larger, smaller) than its parent atom.
6. A positive ion is (larger, smaller) than its parent atom.
7. As you go from left to right across a period, the first ionization energy generally (decreases, increases). Why? ______________________________________
8. As you go down a group, the first ionization energy generally (decreases, increases). Why? ______________________________________
9. Where is the highest electronegativity found? ____________________________
10. Where is the lowest electronegativity found? _____________________________
11. Elements of Group 1 are called ____________________________________.
12. Elements of Group 2 are called ____________________________________.
13. Elements of Group 3–12 are called __________________________________.
14. As you go from left to right across the periodic table, the elements go from (metals, nonmetals) to (metals, nonmetals).
15. Group 17 elements are called ______________________________________.
16. The most active element in Group 17 is ________________________________.
17. Group 18 elements are called ______________________________________.
18. What sublevels are filling across the Transition Elements? __________________
19. Elements within a group have a similar number of _________________________.
20. Elements across a series have the same number of ________________________.
21. A colored ion generally indicates a ___________________________________.
22. As you go down a group, the elements generally become (more, less) metallic.
23. The majority of elements in the periodic table are (metals, nonmetals).
24. Elements in the periodic table are arranged according to their ______________.
25. An element with both metallic and nonmetallic properties is called a __________
    ____

Name________________________________

Periodic Table Puzzle

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
I																	
	F														G	H	
													B				A
C								E				J					

			D											

Place the letter of each of the above elements next to its description.

1. An alkali metal ________
2. An alkaline earth metal ________
3. An inactive gas ________
4. An active nonmetal ________
5. A semimetal ________
6. An inner transition element ________
7. Its most common oxidation state is -2. ________
8. A metal with more than one oxidation state ________
9. A metal with an oxidation number of +3 ________
10. Has oxidation numbers of +1 and -1 ________

Name________________________________

Periodic Table Puzzle

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
A													E				
																D	
		C												G			B
									H		J						
	F																

										I				

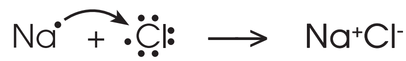
Place the letter of each of the above elements next to its description. Then, using a copy of the periodic table, write the element's name and symbol.

1. An element in Group 17 ______________________________
2. An alkali metal with an oxidation of +1 ______________________________
3. A metalloid ______________________________
4. Has 80 electrons ______________________________
5. Its electron configuration is 2, 4. ______________________________
6. Has 99 protons ______________________________
7. An element in the fifth period with fewer than 50 protons ______________________________
8. Its electron configuration is 2, 8, 18, 18, 8. ______________________________
9. A transition metal with one electron in its outer shell ______________________________
10. An alkaline earth metal ______________________________

Name________________________________

Ionic Bonding

Ionic bonding occurs when a metal transfers one or more electrons to a nonmetal in an effort to attain a stable octet of electrons. For example, the transfer of an electron from sodium to chlorine can be depicted by a Lewis dot diagram.



Calcium would need two chlorine atoms to get rid of its two valence electrons.

Sketch the transfer of electrons in each combination.

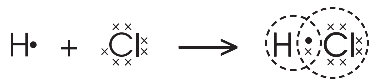
1.	K + F
2.	Mg + I
3.	Be + S
4.	Na + O
5.	Al + Br

Name______________________________

Covalent Bonding

Covalent bonding occurs when two or more nonmetals share electrons, attempting to attain a stable octet of electrons at least part of the time.

Example:



Note that hydrogen is content with 2, not 8, electrons.

Sketch how covalent bonding occurs in each pair of atoms. Atoms may share one, two, or three pairs of electrons.

1. H + H (H_2)
2. F + F (F_2)
3. O + O (O_2)
4. N + N (N_2)
5. C + O (CO_2)
6. H + O (H_2O)

Name________________________________

Types of Chemical Bonds

Identify each compound as *ionic* (metal + nonmetal), *covalent* (nonmetal + nonmetal), or *both* (compound containing a polyatomic ion).

1. $CaCl_2$ ______________________
2. CO_2 ______________________
3. H_2O ______________________
4. $BaSO_4$ ______________________
5. K_2O ______________________
6. NaF ______________________
7. Na_2CO_3 ______________________
8. CH_4 ______________________
9. SO_3 ______________________
10. LiBr ______________________
11. MgO ______________________
12. NH_4Cl ______________________
13. HCl ______________________
14. KI ______________________
15. NaOH ______________________
16. NO_2 ______________________
17. $AlPO_4$ ______________________
18. $FeCl_3$ ______________________
19. P_2O_5 ______________________
20. N_2O_3 ______________________

Name______________________________

Shapes of Molecules

Using VSEPR theory, name and sketch the shape of each molecule.

1. N_2	7. HF
2. H_2O	8. CH_3OH
3. CO_2	9. H_2S
4. NH_3	10. I_2
5. CH_4	11. $CHCl_3$
6. SO_3	12. O_2

Name___________________________________

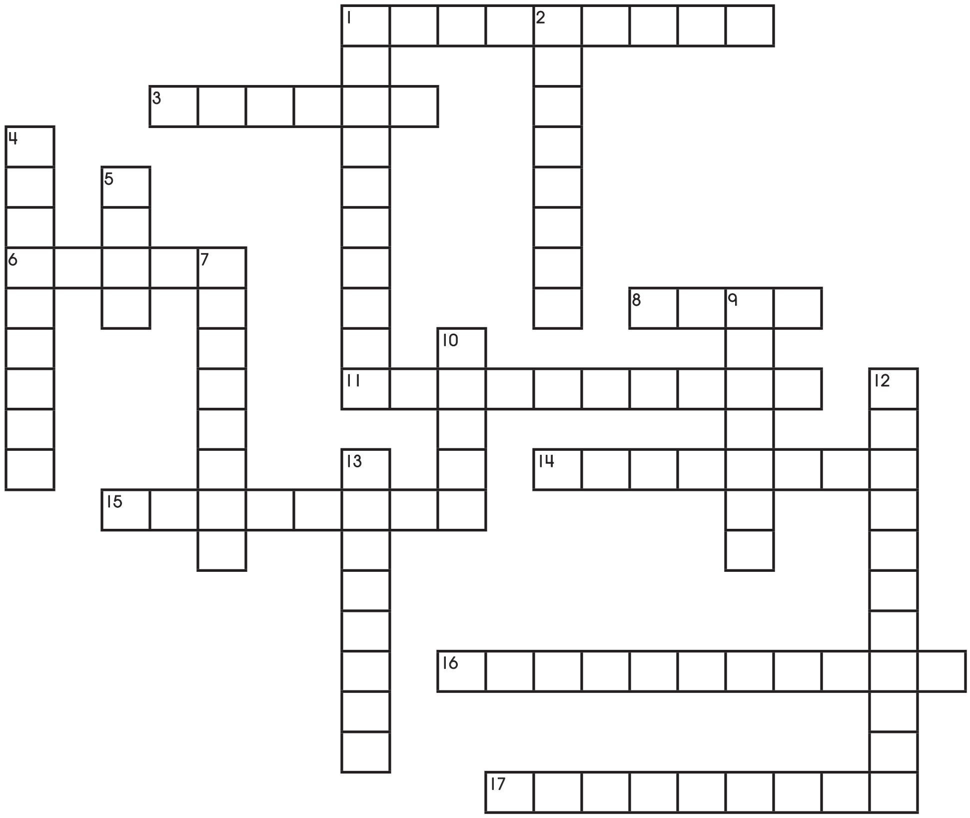
Polarity of Molecules

Identify each molecule as *polar* or *nonpolar*.

1. N_2	7. HF
2. H_2O	8. CH_3OH
3. CO_2	9. H_2S
4. NH_3	10. I_2
5. CH_4	11. $CHCl_3$
6. SO_3	12. O_2

Name________________________________

Chemical Bonding Crossword



Across

1. Ammonia is polar because its shape is ________.
3. Word used to describe a molecule with an unequal charge distribution
6. Type of bond formed between an active metal and a nonmetal
8. The simultaneous attraction of electrons for the nucleii of two or more atoms is a chemical ________.
11. Type of covalent bond in which one atom donates both electrons
14. Bonding that is responsible for the relatively high boiling point of water
15. Type of covalent bond found in diatomic molecules
16. Carbon dioxide is nonpolar because it is ________.
17. Particles formed from covalent bonding

Down

1. Compounds with both ionic and covalent bonds contain this type of ion.
2. Type of bond found in aluminum foil
4. The formulas of ionic compounds must be expressed as ________ formulas.
5. The shape of a water molecule
7. Type of bond found between nonmetals
9. Type of covalent bonding that is found in a diamond
10. Type of covalent bond found between atoms of different electronegativity values
12. Force of attraction between nonpolar molecules
13. Element with the highest electronegativity value

Name________________________________

Writing Formulas (Crisscross Method)

Write the formula of the compound produced from the listed ions.

	Cl^{-}	CO_3^{-2}	OH^{-}	SO_4^{-2}	PO_4^{-3}	NO_3^{-}
Na^{+}						
NH_4^{+}						
K^{+}						
Ca^{2+}						
Mg^{2+}						
Zn^{2+}						
Fe^{3+}						
Al^{3+}						
Co^{3+}						
Fe^{2+}						
H^{+}						

Name________________________________

Naming Ionic Compounds

Name each compound using the Stock Naming System.

1. $CaCO_3$ ____________________
2. KCl ____________________
3. $FeSO_4$ ____________________
4. $LiBr$ ____________________
5. $MgCl_2$ ____________________
6. $FeCl_3$ ____________________
7. $Zn_3(PO_4)_2$ ____________________
8. NH_4NO_3 ____________________
9. $Al(OH)_3$ ____________________
10. $CuC_2H_3O_2$ ____________________
11. $PbSO_3$ ____________________
12. $NaClO_3$ ____________________
13. CaC_2O_4 ____________________
14. Fe_2O_3 ____________________
15. $(NH_4)_3PO_4$ ____________________
16. $NaHSO_4$ ____________________
17. Hg_2Cl_2 ____________________
18. $Mg(NO_2)_2$ ____________________
19. $CuSO_4$ ____________________
20. $NaHCO_3$ ____________________
21. $NiBr_3$ ____________________
22. $Be(NO_3)_2$ ____________________
23. $ZnSO_4$ ____________________
24. $AuCl_3$ ____________________
25. $KMnO_4$ ____________________

Name________________________________

Naming Molecular Compounds

Name each covalent compound.

1. CO_2 ____________________
2. CO ____________________
3. SO_2 ____________________
4. SO_3 ____________________
5. N_2O ____________________
6. NO ____________________
7. N_2O_3 ____________________
8. NO_2 ____________________
9. N_2O_4 ____________________
10. N_2O_5 ____________________
11. PCl_3 ____________________
12. PCl_5 ____________________
13. NH_3 ____________________
14. SCl_6 ____________________
15. P_2O_5 ____________________
16. CCl_4 ____________________
17. SiO_2 ____________________
18. CS_2 ____________________
19. OF_2 ____________________
20. PBr_3 ____________________

Name________________________

Naming Acids

Name each acid.

1. HNO_3 ________________
2. HCl ________________
3. H_2SO_4 ________________
4. H_2SO_3 ________________
5. $HC_2H_3O_2$ ________________
6. HBr ________________
7. HNO_2 ________________
8. H_3PO_4 ________________
9. H_2S ________________
10. H_2CO_3 ________________

Write the formula of each acid.

11. sulfuric acid ________________
12. nitric acid ________________
13. hydrochloric acid ________________
14. acetic acid ________________
15. hydrofluoric acid ________________
16. phosphorous acid ________________
17. carbonic acid ________________
18. nitrous acid ________________
19. phosphoric acid ________________
20. hydrosulfuric acid ________________

Name____________________

Writing Formulas from Names

Write the formula of each compound.

1. ammonium phosphate ____________________
2. iron(II) oxide ____________________
3. iron(III) oxide ____________________
4. carbon monoxide ____________________
5. calcium chloride ____________________
6. potassium nitrate ____________________
7. magnesium hydroxide ____________________
8. aluminum sulfate ____________________
9. copper(II) sulfate ____________________
10. lead(IV) chromate ____________________
11. diphosphorus pentoxide ____________________
12. potassium permanganate ____________________
13. sodium hydrogen carbonate ____________________
14. zinc nitrate ____________________
15. aluminum sulfite ____________________

Name________________________________

Gram Formula Mass

Determine the gram formula mass (the mass of one mole) of each compound.

1. $KMnO_4$ ____________________
2. KCl ____________________
3. Na_2SO_4 ____________________
4. $Ca(NO_3)_2$ ____________________
5. $Al_2(SO_4)_3$ ____________________
6. $(NH_4)_3PO_4$ ____________________
7. $CuSO_4 \bullet 5H_2O$ ____________________
8. $Mg_3(PO_4)_2$ ____________________
9. $Zn(C_2H_3O_2)_2 \bullet 2H_2O$ ____________________
10. $Zn_3(PO_4)_2 \bullet 4H_2O$ ____________________
11. H_2CO_3 ____________________
12. $Hg_2Cr_2O_7$ ____________________
13. $Ba(ClO_3)_2$ ____________________
14. $Fe_2(SO_3)_3$ ____________________
15. $NH_4C_2H_3O_2$ ____________________

Name________________________________

Moles and Mass

Determine the number of moles in each quantity.

1.	25 g of $NaCl$
2.	125 g of H_2SO_4
3.	100. g of $KMnO_4$
4.	74 g of KCl
5.	35 g of $CuSO_4 \cdot 5H_2O$

Determine the number of grams in each quantity.

6.	2.5 mol of $NaCl$
7.	0.50 mol of H_2SO_4
8.	1.70 mol of $KMnO_4$
9.	0.25 mol of KCl
10.	3.2 mol of $CuSO_4 \cdot 5H_2O$

Name________________________________

The Mole and Volume

For gases at STP (273 K and 1 atm pressure), one mole occupies a volume of 22.4 L. Identify the volume each quantity of gas will occupy at STP.

1.	1.00 mole of H_2
2.	3.20 moles of O_2
3.	0.750 mole of N_2
4.	1.75 moles of CO_2
5.	0.50 mole of NH_3
6.	5.0 g of H2
7.	100. g of O_2
8.	28.0 g of N_2
9.	60. g of CO_2
10.	10. g of NH_3

Name________________________________

The Mole and Avogadro's Number

One mole of a substance contains Avogadro's number (6.02×10^{23}) of molecules.

Identify how many molecules are in each quantity.

1.	2.0 mol
2.	1.5 mol
3.	0.75 mol
4.	15 mol
5.	0.35 mol

Identify how many moles are in the number of molecules listed.

6.	6.02×10^{23}
7.	1.204×10^{24}
8.	1.5×10^{20}
9.	3.4×10^{26}
10.	7.5×10^{19}

Name________________________________

Mixed Mole Problems

Solve each problem.

1. How many grams are there in 1.5×10^{25} molecules of CO_2?

2. What volume would the CO_2 in problem 1 occupy at STP?

3. A sample of NH_3 gas occupies 75.0 liters at STP. How many molecules is this?

4. What is the mass of the sample of NH_3 in problem 3?

5. How many atoms are there in 1.3×10^{22} molecules of NO_2?

6. A 5.0 g sample of O_2 is in a container at STP. What volume is the container?

7. How many molecules of O_2 are in the container in problem 6? How many atoms of oxygen?

Name______________________________

Percentage Composition

Determine the percentage composition of each compound.

1. $KMnO_4$

   K = ______________

   Mn = ______________

   O = ______________

---

2. HCl

   H = ______________

   Cl = ______________

---

3. $Mg(NO_3)_2$

   Mg = ______________

   N = ______________

   O = ______________

---

4. $(NH_4)PO_4$

   N = ______________

   H = ______________

   P = ______________

   O = ______________

---

5. $Al_2(SO_4)_3$

   Al = ______________

   S = ______________

   O = ______________

Solve each problem.

6. How many grams of oxygen can be produced from the decomposition of 100. g of $KClO_3$? ____________________

7. How much iron can be recovered from 25.0 g of Fe_2O_3? ____________________

8. How much silver can be produced from 125 g of Ag_2S? ____________________

Name_______________________________

Determining Empirical Formulas

Identify the empirical formula (lowest whole number ratio) of each compound.

1. 75% carbon, 25% hydrogen

2. 52.7% potassium, 47.3% chlorine

3. 22.1% aluminum, 25.4% phosphorus, 52.5% oxygen

4. 13% magnesium, 87% bromine

5. 32.4% sodium, 22.5% sulfur, 45.1% oxygen

6. 25.3% copper, 12.9% sulfur, 25.7% oxygen, 36.1% water

Name_______________________________

Determining Molecular Formulas (True Formulas)

Solve each problem.

1.	The empirical formula of a compound is NO_2. Its molecular mass is 92 g/mol. What is its molecular formula?
2.	The empirical formula of a compound is CH_2. Its molecular mass is 70 g/mol. What is its molecular formula?
3.	A compound is found to be 40.0% carbon, 6.7% hydrogen and 53.5% oxygen. Its molecular mass is 60. g/mol. What is its molecular formula?
4.	A compound is 64.9% carbon, 13.5% hydrogen, and 21.6% oxygen. Its molecular mass is 74 g/mol. What is its molecular formula?
5.	A compound is 54.5% carbon, 9.1% hydrogen, and 36.4% oxygen. Its molecular mass is 88 g/mol. What is its molecular formula?

Name________________________________

Composition of Hydrates

A **hydrate** is an ionic compound with water molecules loosely bonded to its crystal structure. The water is in a specific ratio to each formula unit of the salt. For example, the formula $CuSO_4 \bullet 5H_20$ indicates that there are five water molecules for every one formula unit of $CuSO_4$.

Solve each problem.

1.	What percentage of water is found in $CuSO_4 \bullet 5H_2O$?
2.	What percentage of water is found in $Na_2S \bullet 9H_2O$?
3.	A 5.0 g sample of a hydrate of $BaCl_2$ was heated, and only 4.3 g of the anhydrous salt remained. What percentage of water was in the hydrate?
4.	A 2.5 g sample of a hydrate of $Ca(NO_3)_2$ was heated, and only 1.7 g of the anhydrous salt remained. What percentage of water was in the hydrate?
5.	A 3.0 g sample of $Na_2CO_3 \bullet H_2O$ is heated to constant mass. How much anhydrous salt remains?
6.	A 5.0 g sample of $Cu(NO_3)_2 \bullet nH_2O$ is heated to constant mass. How much anhydrous salt remains?

Name________________________________

Balancing Chemical Equations

Rewrite and balance each equation.

1. $N_2 + H_2 \rightarrow NH_3$ ________________________________
2. $KClO_3 \rightarrow KCl + O_2$ ________________________________
3. $NaCl + F_2 \rightarrow NaF + Cl_2$ ________________________________
4. $H_2 + O_2 \rightarrow H_2O$ ________________________________
5. $AgNO_3 + MgCl_2 \rightarrow AgCl + Mg(NO_3)_2$ ________________________________
6. $AlBr_3 + K_2SO \rightarrow KBr + Al_2(SO_4)_3$ ________________________________
7. $CH_4 + O_2 \rightarrow CO_2 + H_2O$ ________________________________
8. $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$ ________________________________
9. $C_8H_{18} + O_2 \rightarrow CO_2 + H_2O$ ________________________________
10. $FeCl_3 + NaOH \rightarrow Fe(OH)_3 + NaCl$ ________________________________
11. $P + O_2 \rightarrow P_2O_5$ ________________________________
12. $Na + H_2O \rightarrow NaOH + H_2$ ________________________________
13. $Ag_2O \rightarrow Ag + O_2$ ________________________________
14. $S_8 + O_2 \rightarrow SO_3$ ________________________________
15. $CO_2 + H_2O \rightarrow C_6H_{12}O_6 + O_2$ ________________________________
16. $K + MgBr_2 \rightarrow KBr + Mg$ ________________________________
17. $HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$ ________________________________

Name________________________________

Word Equations

Rewrite each word equation as a chemical equation. Then, balance the equation.

1.	zinc + lead(II) nitrate yield zinc nitrate + lead
2.	aluminum bromide + chlorine yield aluminum chloride + bromine
3.	sodium phosphate + calcium chloride yield calcium + sodium chloride
4.	potassium chlorate, when heated, yields potassium chloride + oxygen gas
5.	aluminum + hydrochloric acid yield aluminum chloride + hydrogen gas
6.	calcium hydroxide + phosphoric acid yield calcium phosphate + water
7 .	copper + sulfuric acid yield copper(II) sulfate + water + sulfur dioxide
8.	hydrogen + nitrogen monoxide yield water + nitrogen

Name________________________________

Classification of Chemical Reactions

Identify each reaction as *synthesis, decomposition, cationic* or *anionic single replacement*, or *double replacement*.

1. $2H_2 + O_2 \rightarrow 2H_2O$

   ____

2. $2H_2O \rightarrow 2H_2 + O_2$

   ____

3. $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

   ____

4. $2CO + O_2 \rightarrow 2CO_2$

   ____

5. $2HgO \rightarrow 2Hg + O_2$

   ____

6. $2KBr + Cl_2 \rightarrow 2KCl + Br_2$

   ____

7. $CaO + H_2O \rightarrow Ca(OH)_2$

   ____

8. $AgNO + NaCl \rightarrow AgCl + NaNO_3$

   ____

9. $2H_2O_2 \rightarrow 2H_2O + O_2$

   ____

10. $Ca(OH)_2 + H_2SO_4 \rightarrow CaSO_4 + 2H_2O$

   ____

Name____________________________________

Predicting Products of Chemical Reactions

Predict the product in each reaction. Then, write the balanced equation and classify the reaction.

1.	magnesium bromide + chlorine
2.	aluminum + iron(III) oxide
3.	silver nitrate + zinc chloride
4.	hydrogen peroxide (catalyzed by manganese dioxide)
5.	zinc + hydrochloric acid
6.	sulfuric acid + sodium hydroxide
7.	sodium + hydrogen
8.	acetic acid + copper

Name_______________________________

Stoichiometry: Mole-Mole Problems

Solve each problem.

1.	$N_2 + 3H_2 \rightarrow 2NH_3$ How many moles of hydrogen are needed to completely react with two moles of nitrogen?
2.	$2KClO_3 \rightarrow 2KCl + 3O_2$ How many moles of oxygen are produced by the decomposition of six moles of potassium chlorate?
3.	$Zn + 2HCl \rightarrow ZnCl_2 + H_2$ How many moles of hydrogen are produced from the reaction of three moles of zinc with an excess of hydrochloric acid?
4.	$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$ How many moles of oxygen are necessary to react completely with four moles of propane (C_3H_8)?
5.	$K_3PO_4 + Al(NO_3)_3 \rightarrow 3KNO_3 + AlPO_4$ How many moles of potassium nitrate (KNO_3) are produced when six moles of potassium phosphate (KPO_4) react with two moles of aluminum nitrate($Al(NO_3)_3$)?

Name________________________________

Stoichiometry: Volume-Volume Problems

Solve each problem.

1. $N_2 + 3H_2 \rightarrow 2NH_3$
   What volume of hydrogen is necessary to react with five liters of nitrogen to produce ammonia? (Assume constant temperature and pressure.)

2. What volume of ammonia is produced in the reaction in problem 1?

3. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
   If 20 liters of oxygen are consumed in the above reaction, how many liters of carbon dioxide are produced?

4. $2H_2O \rightarrow 2H + O_2$
   If 30 mL of hydrogen are produced in the above reaction, how many milliliters of oxygen are produced?

5. $2CO + O_2 \rightarrow 2CO_2$
   How many liters of carbon dioxide are produced if 75 liters of carbon monoxide are burned in oxygen? How many liters of oxygen are necessary?

Name____________________________

Stoichiometry: Mass-Mass Problems

Solve each problem.

1.	$2KClO_3 \rightarrow 2KCl + 3O_2$ How many grams of potassium chloride are produced if 25 g of potassium chlorate decompose?
2.	$N_2 + 3H_2 \rightarrow 2NH_3$ How many grams of hydrogen are necessary to react completely with 50.0 g of nitrogen in the above reaction?
3.	How many grams of ammonia are produced in the reaction in problem 2?
4.	$2AgNO_3 + BaCl \rightarrow 2AgCl + Ba(NO_3)$ How many grams of silver chloride are produced from 5.0 g of silver nitrate reacting with an excess of barium chloride?
5.	How much barium chloride is necessary to react with the silver nitrate in problem 4?

Name____________________________

Stoichiometry: Mixed Problems

Solve each problem.

1.	$N_2 + 3H_2 \rightarrow 2NH_3$ What volume of NH_3 at STP is produced if 25.0 g of N_2 is reacted with an excess of H_2?
2.	$2KClO_3 \rightarrow 2KCl + 3O_2$ If 5.0 g of $KClO_3$ are decomposed, what volume of O_2 is produced at STP?
3.	How many grams of KCl are produced in problem 2?
4.	$Zn + 2HCl \rightarrow ZnCl_2 + H_2$ What volume of hydrogen at STP is produced when 2.5 g of zinc react with an excess of hydrochloric acid?
5.	$H_2SO_4 + 2NaOH \rightarrow H_2O + Na_2SO_4$ How many molecules of water are produced if 2.0 g of sodium sulfate are produced in the above reaction?
6.	$2AlCl_3 \rightarrow 2Al + 3Cl_2$ If 10.0 g of aluminum chloride are decomposed, how many molecules of Cl_2 are produced?

Name____________________________

Stoichiometry: Limiting Reagent

Solve each problem.

1.	$N_2 + 3H_2 \longrightarrow 2NH_3$ How many grams of NH_3 can be produced from the reaction of 28 g of N_2 and 25 g of H_2?
2.	How much of the excess reagent in problem 1 is left over?
3.	$Mg + 2HCl \longrightarrow MgCl_2 + H_2$ What volume of hydrogen at STP is produced from the reaction of 50.0 g of Mg and the equivalent of 75 g of HCl?
4.	How much of the excess reagent in problem 3 is left over?
5.	$3AgNO_3 + Na_3PO_4 \longrightarrow Ag_3PO_4 + 3NaNO_3$ Silver nitrate and sodium phosphate are reacted in equal amounts of 200. g each. How many grams of silver phosphate are produced?
6.	How much of the excess reagent in problem 5 is left over?

Name________________________________

Solubility Curves

Answer each question based on the solubility curve shown.

1. Which salt is least soluble in water at 20°C? ______________

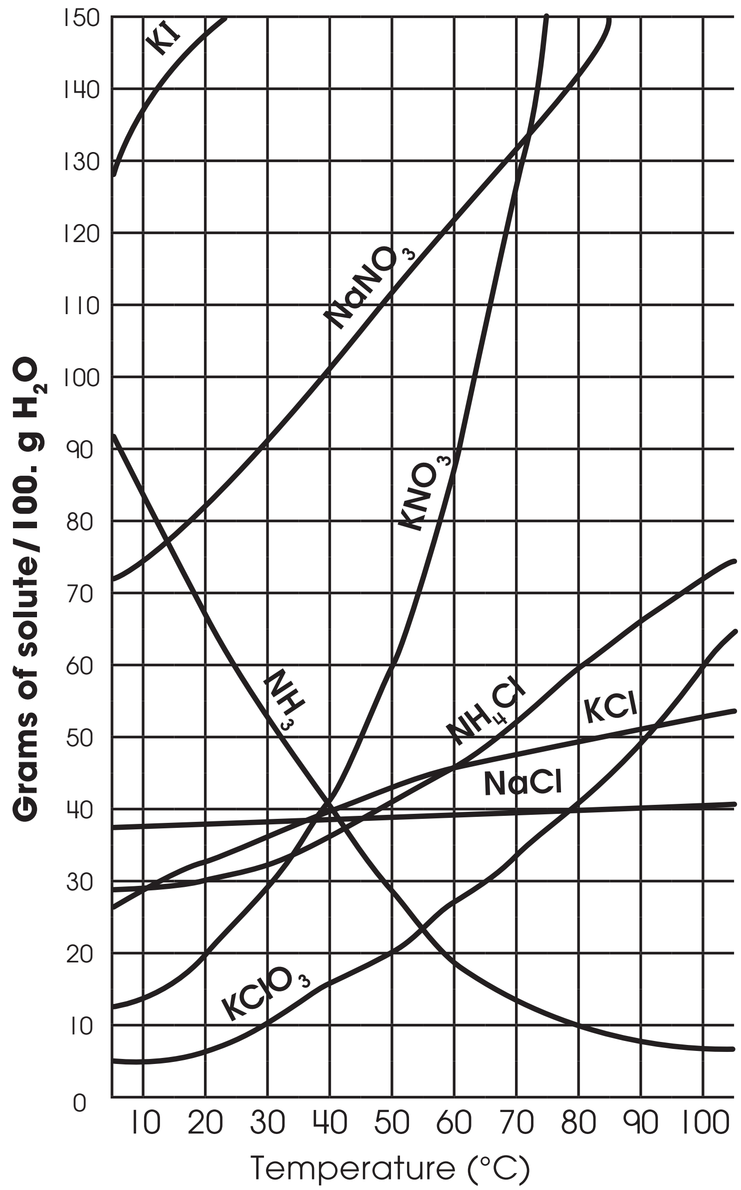
2. How many grams of potassium chloride can be dissolved in 200 g of water at 80°C? ______________

3. At 40°C, how much potassium nitrate can be dissolved in 300 g of water? ______________

4. Which salt shows the least change in solubility from 0°C to 100°C? ______________

5. At 30°C, 85 g of sodium nitrate are dissolved in 100 g of water. Is this solution *saturated*, *unsaturated*, or *supersaturated*? ______________

6. A saturated solution of potassium chlorate is formed from 100 g of water. If the saturated solution is cooled from 80°C to 50°C, how many grams of precipitate are formed? ______________



7. What compound shows a decrease in solubility from 0°C to 100°C? ______________

8. Which salt is most soluble at 10°C? ______________

9. Which salt is least soluble at 50°C? ______________

10. Which salt is least soluble at 90°C? ______________

Name________________________________

Molarity (M)

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liter of solution}}$$

Solve each problem.

1.	What is the molarity of a solution in which 58 g of NaCl are dissolved in 1.0 L of solution?
2.	What is the molarity of a solution in which 10.0 g of $AgNO_3$ are dissolved in 500. mL of solution?
3.	How many grams of KNO_3 should be used to prepare 2.00 L of a 0.500 M solution?
4.	To what volume should 5.0 g of KCl be diluted in order to prepare a 0.25 M solution?
5.	How many grams of $CuSO_4{\cdot}5H_2O$ are needed to prepare 100. mL of a 0.10 M solution?

Name____________________________

Molarity by Dilution

Acids are usually acquired from chemical supply houses in concentrated form. These acids are diluted to the desired concentration by adding water. Since moles of acid before dilution equal moles of acid after dilution, and moles of acid = $M \times V$, then $M_1 \times V_1 = M_2 \times V_2$.

Solve each problem.

1. How much concentrated 18 M sulfuric acid is needed to prepare 250 mL of a 6.0 M solution?

2. How much concentrated 12 M hydrochloric acid is needed to prepare $1\bar{0}0$ mL of a 2.0 M solution?

3. To what volume should 25 mL of 15 M nitric acid be diluted to prepare a 3.0 M solution?

4. How much water should be added to 50. mL of 12 M hydrochloric acid to produce a 4.0 M solution?

5. How much water should be added to100. mL of 18 M sulfuric acid to prepare a 1.5 M solution?

Name________________________________

Molality (m)

$$\text{Molality} = \frac{\text{moles of solute}}{\text{kg of solvent}}$$

Solve each problem.

1.	What is the molality of a solution in which 3.0 moles of NaCl are dissolved in 1.5 kg of water?
2.	What is the molality of a solution in which 25 g of NaCl are dissolved in 2.0 kg of water?
3.	What is the molality of a solution in which 15 g of I_2 are dissolved in 500. g of alcohol?
4.	How many grams of I_2 should be added to 750 g of CCl_4 to prepare a 0.020 m solution?
5.	How much water should be added to 5.00 g of KCl to prepare a 0.500 m solution?

Name________________________________

Normality (N)

normality = molarity × total positive oxidation number of solute

Example: What is the normality of 3.0 M of H_2SO_4?

Since the total positive oxidation number of H_2SO_4 is +2 (2 H^+), N = 6.0.

Solve each problem.

1.	What is the normality of a 2.0 M NaOH solution?
2.	What is the normality of a 2.0 M H_3PO_4 solution?
3.	A solution of H_2SO_4 is 3.0 N. What is its molarity?
4.	What is the normality of a solution in which 2.0 g of $Ca(OH)_2$ is dissolved in 1.0 L of solution?
5.	How much $AlCl_3$ should be dissolved in 2.00 L of solution to produce a 0.150 N solution?

Name________________________________

Electrolytes

Electrolytes are substances that break up (dissociate or ionize) in water to produce ions. These ions are capable of conducting an electric current.

Generally, electrolytes consist of acids, bases, and salts (ionic compounds). Nonelectrolytes are usually covalent compounds, with the exception of acids.

Check the appropriate column to classify each compound as either an electrolyte or a nonelectrolyte.

Compound	Electrolyte	Nonelectrolyte
1. NaCl		
2. CH_3OH (methyl alcohol)		
3. $C_3H_5(OH)_3$ (glycerol)		
4. HCl		
5. $C_6H_{12}O_6$ (sugar)		
6. NaOH		
7. C_2H_5OH (ethyl alcohol)		
8. CH_3COOH (acetic acid)		
9. NH_4OH ($NH_3 + H_2O$)		
10. H_2SO_4		

Name______________________________

Effect of a Solute on Freezing and Boiling Points

Use the following formulas to calculate changes in freezing and boiling points due to the presence of a nonvolatile solute. The freezing point is always lowered; the boiling point is always raised.

$\Delta T_F = M \times d.f. \times k_F$ $\quad$ $k_B H_2O = 0.52°C/M$

$\Delta T_B = M \times d.f. \times k_B$ $\quad$ $k_F H_2O = 1.86°C/M$

M = molality of solution

k_F and k_B = constants for particular solvent

d.f. = dissociation factor (how many particles the solute breaks up into; for a nonelectrolyte, *d.f.* = 1)

(The theoretical dissociation factor is always greater than observed effect.)

Solve each problem.

1. What is the new boiling point if 25 g of NaCl are dissolved in 1.0 kg of water?

2. What is the freezing point of the solution in problem 1?

3. What are the new freezing and boiling points of water if 50. g of ethylene glycol (molecular mass = 62 g/M) are added to 50. g of water?

4. When 5.0 g of a nonelectrolyte are added to 25 g of water, the new freezing point is -2.5°C. What is the molecular mass of the unknown compound?

Name________________________________

Solubility (Polar vs. Nonpolar)

Generally, "like dissolves like." Polar molecules dissolve other polar molecules and ionic compounds. Nonpolar molecules dissolve other nonpolar molecules. Alcohols, which have characteristics of both, tend to dissolve in both types of solvents but will not dissolve ionic solids.

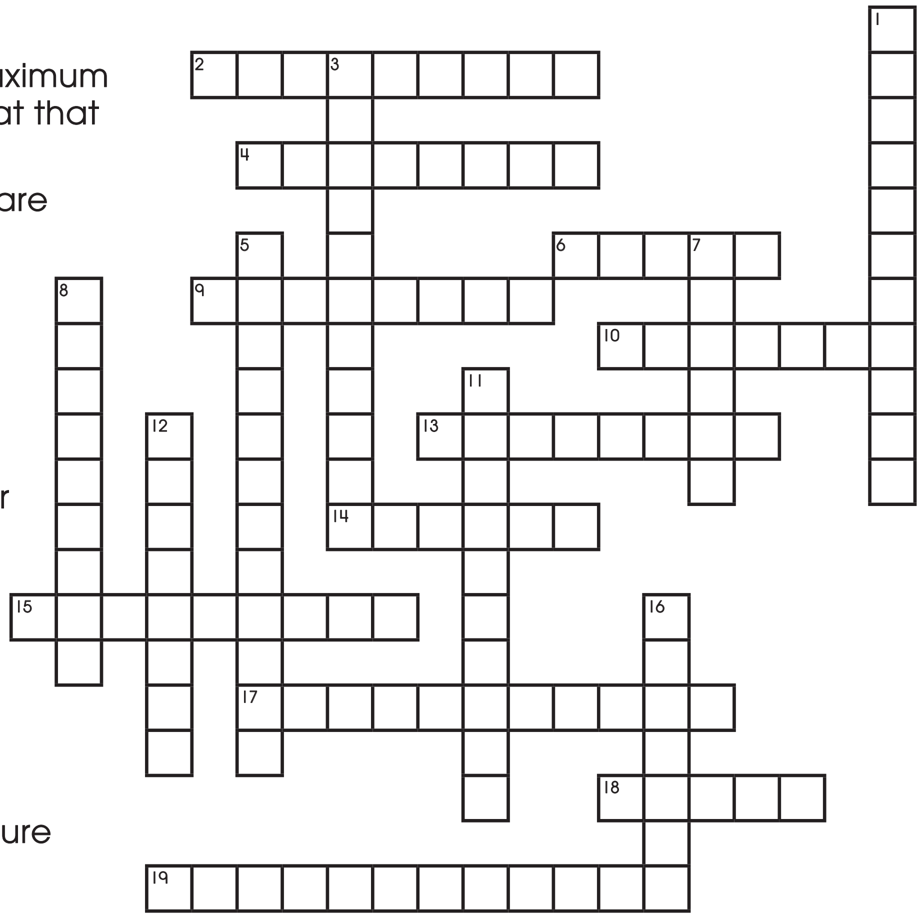
Check the appropriate columns to indicate whether the solute is soluble in a polar or nonpolar solvent.

Solutes	**Solvents** water	CCl_4	alcohol
1. NaCl			
2. I_2			
3. ethanol			
4. benzene			
5. Br_2			
6. KNO_3			
7. toluene			
8. $Ca(OH)_2$			

Name________________________________

Solutions Crossword

Across



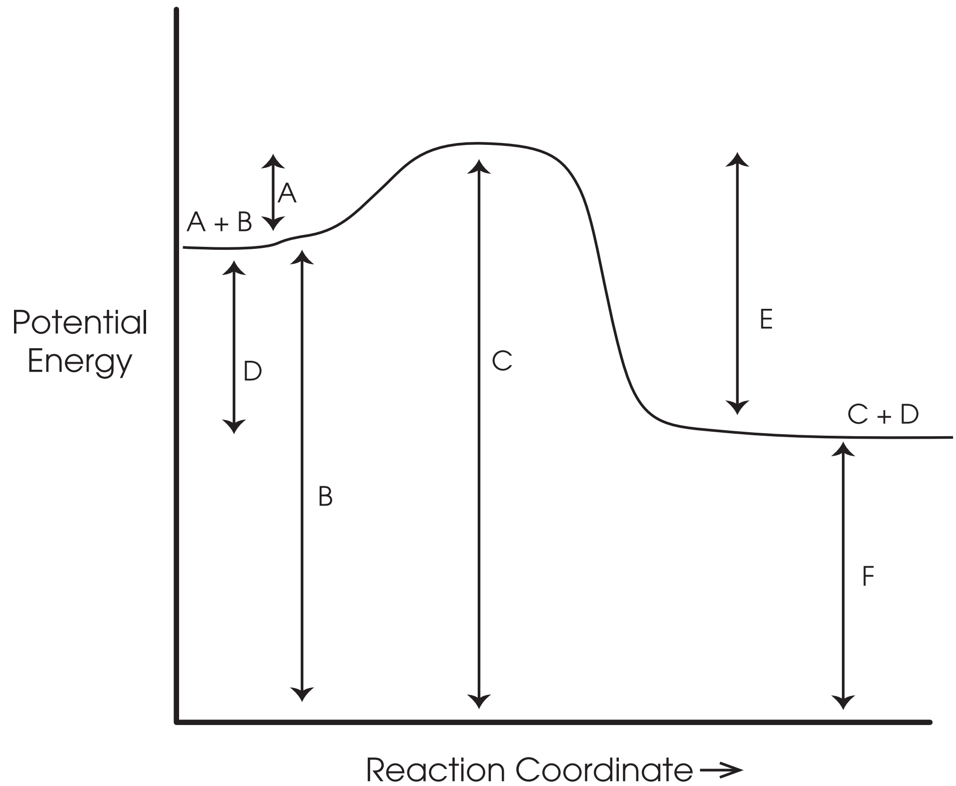
2. Solution containing the maximum amount of solute possible at that temperature
4. Two liquids which can mix are said to be _____________.
6. The presence of a nonvolatile solute will _____________ the boiling point of a solvent.
9. A homogeneous mixture
10. Substance present in larger amounts in a mixture
13. Moles of a solute per kilogram of solvent
14. Solution containing a relatively large amount of solvent
15. The solubility of gases _____________ as temperature increases.
17. State in which the rate of dissolving is equal to the rate of precipitation
18. The presence of a nonvolatile solute will _____________ the freezing point of a solvent.
19. These substances dissociate or ionize in water and are then able to conduct an electric current.

Down

1. Properties that depend on the number of particles in a solution
3. Solution in which more solute can be dissolved
5. Solution containing a relatively large amount of dissolved solute
7. Substance present in a smaller amount in a mixture
8. The solubility of most solids _____________ as temperature increases.
11. Maximum amount of solute that can dissolve in a stated amount of solute at a given temperature
12. Moles of solute per liter of solution
16. Solutions in which water is the solvent are called _____________.

Name________________________________

Potential Energy Diagram



A + B ⟷ C + D + energy

Answer each question using the graph shown.

1. Is the above reaction endothermic or exothermic? ______________________
2. Which letter represents the potential energy of the reactants? ___________
3. Which letter represents the potential energy of the products? ___________
4. Which letter represents the heat of reaction (ΔH)? ___________
5. Which letter represents the activation energy of the forward reaction? ___________
6. Which letter represents the activation energy of the reverse reaction? ___________
7. Which letter represents the potential energy of the activated complex? ___________
8. Is the reverse reaction endothermic or exothermic? ______________________
9. If a catalyst were added, what letter(s) would change? ______________________

Name________________________________

Entropy

Entropy is the degree of randomness in a substance. The symbol for change in entropy is ΔS.

Solids are very ordered and have low entropy. Liquids and aqueous ions have more entropy because they move about more freely. Gases have an even larger amount of entropy. According to the Second Law of Thermodynamics, nature is always proceeding to a state of higher entropy.

Determine whether each reaction shows an *increase* or *decrease* in entropy.

1. $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$ ____________________
2. $H_2O(l) \rightarrow H_2O(s)$ ____________________
3. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ ____________________
4. $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ ____________________
5. $KCl(s) \rightarrow KCl(l)$ ____________________
6. $CO_2(s) \rightarrow CO_2(g)$ ____________________
7. $H^+(aq) + C_2H_3O_2^-(aq) \rightarrow HC_2H_3O_3(l)$ ____________________
8. $C(s) + O_2(g) \rightarrow CO_2(g)$ ____________________
9. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ ____________________
10. $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$ ____________________
11. $2N_2O_5(g) \rightarrow 4NO_2(g) + O_2(g)$ ____________________
12. $2Al(s) + 3I_2(s) \rightarrow 2AlI_3(s)$ ____________________
13. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ ____________________
14. $2NO(g) \rightarrow N_2(g) + O_2(g)$ ____________________
15. $H_2O(g) \rightarrow H_2O(l)$ ____________________

Name______________________________

Gibbs Free Energy

For a reaction to be spontaneous, the sign of ΔG (Gibbs free energy) must be negative. The mathematical formula for this value is:

$$\Delta G = \Delta H - T\Delta S$$

where ΔH = change in enthalpy or heat of reaction
T = temperature in Kelvin
ΔS = change in entropy or randomness

Complete the table for the sign of ΔG: +, –, or *undetermined*. When conditions allow for an undetermined sign of ΔG, temperature will decide spontaneity.

ΔH	ΔS	ΔG
–	+	
+	–	
–	–	
+	+	

Answer each question.

1. The conditions in which ΔG is always negative are when ΔH is ______________________ and ΔS is ______________________.
2. The conditions in which ΔG is always positive are when ΔH is ______________________ and ΔS is ______________________.
3. When the situation is indeterminate, a low temperature favors the (entropy, enthalpy) factor and a high temperature favors the (entropy, enthalpy) factor.

Answer problems 4–6 with *always*, *sometimes*, or *never*.

4. The reaction: $Na(OH)_5 \rightarrow Na^+(aq) + OH^-(aq)$ + energy will ______________________ be spontaneous.
5. The reaction: energy + $2H_2(g) + O_2(g) + 2H_2O(l)$ will ______________________ be spontaneous.
6. The reaction: energy + $H_2O(s) \rightarrow H_2O(l)$ will ______________________ be spontaneous.
7. What is the value of ΔG if ΔH = 32.0 kJ, ΔS = +25.0 kJ/K and T = 293 K? ____________
8. Is the reaction in problem 7 spontaneous? ______________________
9. What is the value of ΔG if ΔH = +12.0 kJ, ΔS = 5.00 kJ/K and T = 290. K? ____________
10. Is the reaction in problem 9 spontaneous? ______________________

Name________________________________

Equilibrium Constant (K)

Write the expression for the equilibrium constant (K) for each reaction.

1. $N_2(g) + 3H_2(g) \longleftrightarrow 2NH_3(g)$
2. $2KClO_3(s) \longleftrightarrow 2KCl(s) + 3O_2(g)$
3. $H_2O(l) \longleftrightarrow H^+(aq) + OH^-(aq)$
4. $2CO(g) + O_2(g) \longleftrightarrow 2CO_2(g)$
5. $Li_2CO_3(s) \rightarrow 2Li^+(aq) + CO_3^{2-}(aq)$

Name________________________________

Calculations Using the Equilibrium Constant

Using the equilibrium constant expressions you determined on page 80, calculate the value of K when:

1. $(NH_3) = 0.0100$ M, $(N_2) = 0.0200$ M, $(H_2) = 0.0200$ M
2. $(O_2) = 0.0500$ M
3. $(H^+) = 1 \times 10^{-8}$ M, $(OH^-) = 1 \times 10^{-6}$ M
4. $(CO) = 2.0$ M, $(O_2) = 1.5$ M, $(CO_2) = 3.0$ M
5. $(Li^+) = 0.2$ M, $(CO_3^{-2}) = 0.1$ M

Name________________________________

Le Chatelier's Principle

Le Chatelier's principle states that when a system at equilibrium is subjected to a stress, the system will shift its equilibrium point in order to relieve the stress.

Complete the chart by writing *left, right,* or *none* for equilibrium shift. Then, write *decreases, increases*, or *remains the same* for the concentrations of reactants and products, and for the value of K. The first one has been done for you.

$$N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g) + 22.0 \text{ kcal}$$

Stress	Equilibrium Shift	(N_2)	(H_2)	(NH_3)	K
1. Add N_2	right	———	decreases	increases	remains the same
2. Add H_2			———		
3. Add NH_3				———	
4. Remove N_2		———			
5. Remove H_2			———		
6. Remove NH_3				———	
7. Increase temperature					
8. Decrease temperature					
9. Increase pressure					
10. Decrease pressure					

Name________________________________

Le Chatelier's Principle (Cont.)

$$12.6 \text{ kcal} + H_2(g) + I_2(g) \leftrightarrow 2HI(g)$$

Stress	Equilibrium Shift	(H_2)	(I_2)	(HI)	K
11. Add H_2	right	———	decreases	increases	remains the same
12. Add I_2			———		
13. Add HI				———	
14. Remove H_2		———			
15. Remove I_2			———		
16. Remove HI				———	
17. Increase temperature					
18. Decrease temperature					
19. Increase pressure					
20. Decrease pressure					

$$NAOH(s) \leftrightarrow Na^+(aq) + OH^-(aq) + 10.6 \text{ kcal}$$

Stress	Equilibrium Shift	Amount NaOH(s)	(Na^+)	(OH^-)	K
21. Add NaOH(s)		———			
22. Add NaCl (Adds Na^+)			———		
23. Add KOH (Adds OH^-)				———	
24. Add H^+ (Removes OH^-)				———	
25. Increase temperature					
26. Decrease temperature					
27. Increase pressure					
28. Decrease pressure					

Name________________________________

Bronsted–Lowry Acids and Bases

According to **Bronsted–Lowry theory**, an acid is a proton (H^+) donor and a base is a proton acceptor.

H^+

Example: $HCl + OH^- \rightarrow Cl^- + H_2O$

The HCl acts as an acid, and the OH^- acts as a base. This reaction is reversible in that the H_2O can give back the proton to the Cl^-.

Label the Bronsted-Lowry acids and bases in each reaction and show the direction of proton transfer.

H^+ **H^+**

Example: $$\underset{\text{acid}}{\mathbf{H_2O}} + \underset{\text{base}}{\mathbf{Cl^-}} \leftrightarrow \underset{\text{base}}{\mathbf{OH^-}} + \underset{\text{acid}}{\mathbf{HCl}}$$

1. $H_2O + H_2O \leftrightarrow H_3O^+ + OH^-$

4. $OH^- + H_3O+ \leftrightarrow H_2O + H_2O$

2. $H_2SO_4 + OH^- \leftrightarrow HSO_4^- + H_2O$

5. $NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$

3. $HSO_4^- + H_2O \leftrightarrow SO_4^{2-} + H_3O^+$

Name____________________________________

Conjugate Acid–Base Pairs

In the exercise on page 84, it was shown that after an acid has given up its proton, it is capable of getting the proton back and acting as a base. **Conjugate base** is what is left after an acid gives up a proton. The stronger the acid, the weaker the conjugate base. The weaker the acid, the stronger the conjugate base.

Complete the chart.

Conjugate Pairs

	Acid	Base	Equation
1.	H_2SO_4	HSO_4^-	$H_2SO_4 \longleftrightarrow H^+ + HSO_4^-$
2.	H_3PO_4		
3.		F^-	
4.		NO_3^-	
5.	$H_2PO_4^-$		
6.	H_2O		
7.		SO_4^{2-}	
8.	HPO_4^{-2}		
9.	NH_4^+		
10.		H_2O	

11. Which is a stronger base, HSO_4^- or $H_2PO_4^-$? ______________________________

12. Which is a weaker base, Cl^- or NO_2^- ? ______________________________

Name________________________________

pH and pOH

The pH of a solution indicates how acidic or basic that solution is. A pH of less than 7 is acidic; a pH of 7 is neutral; and a pH greater than 7 is basic.

Since $(H^+)(OH^-) = 10^{-14}$ at 25°C, if (H^+) is known, the (OH^-) can be calculated and vice versa.

$pH = -\log(H^+)$ So if $(H^+) = 10^{-6}$ M, pH = 6.

$pOH = -\log(OH^-)$ So if $(OH^-) = 10^{-8}$ M, pOH = 8.

Together, pH + pOH = 14.

Complete the chart.

	(H^+)	**pH**	**(OH^-)**	**pOH**	**Acidic or Basic**
1.	10^{-5} M	5	10^{-9} M	9	acidic
2.		7			
3.			10^{-4} M		
4.	10^{-2} M				
5.				11	
6.		12			
7.			10^{-5} M		
8.	10^{-11} M				
9.				13	
10.		6			

Name________________________________

pH of Solutions

Calculate the pH of each solution.

1.	0.01 M HCl
2.	0.0010 M $NaOH$
3.	0.050 M $Ca(OH)_2$
4.	0.030 M HBr
5.	0.150 M KOH
6.	2.0 M $HC_2H_3O_2$ (Assume 5.0% dissociation.)
7.	3.0 M HF (Assume 10.0% dissociation.)
8.	0.50 M HNO_3
9.	2.50 M NH_4OH (Assume 5.00% dissociation.)
10.	5.0 M HNO_2 (Assume 1.0% dissociation.)

Name________________________________

Acid–Base Titration

To determine the concentration of an acid (or base), we can react it with a base (or acid) of known concentration until it is completely neutralized. This point of exact neutralization, known as the **endpoint**, is noted by the change in color of the indicator.

Use the following equation:

$$N_A \times V_A = N_B \times V_B$$

where N = normality
V = volume

Solve each problem.

1. A 25.0 mL sample of HCl was titrated to the endpoint with 15.0 mL of 2.0 N NaOH. What was the normality of the HCl?

2. A 10.0 mL sample of H_2SO_4 was exactly neutralized by 13.5 mL of 1.0 M KOH. What is the normality of the H_2SO_4?

3. How much 1.5 M NaOH is necessary to exactly neutralize 20.0 mL of 2.5 M H_3PO_4?

4. How much 0.5 M HNO_3 is necessary to titrate 25.0 mL of 0.05 M $Ca(OH)_2$ solution to the endpoint?

5. What is the molarity of a NaOH solution if 15.0 mL is exactly neutralized by 7.5 mL of a 0.02 M $HC_2H_3O_2$ solution?

Name________________________________

Hydrolysis of Salts

Salt solutions may be acidic, basic, or neutral, depending on the original acid and base that formed the salt.

strong acid + strong base → neutral salt

strong acid + weak base → acidic salt

weak acid + strong base → basic salt

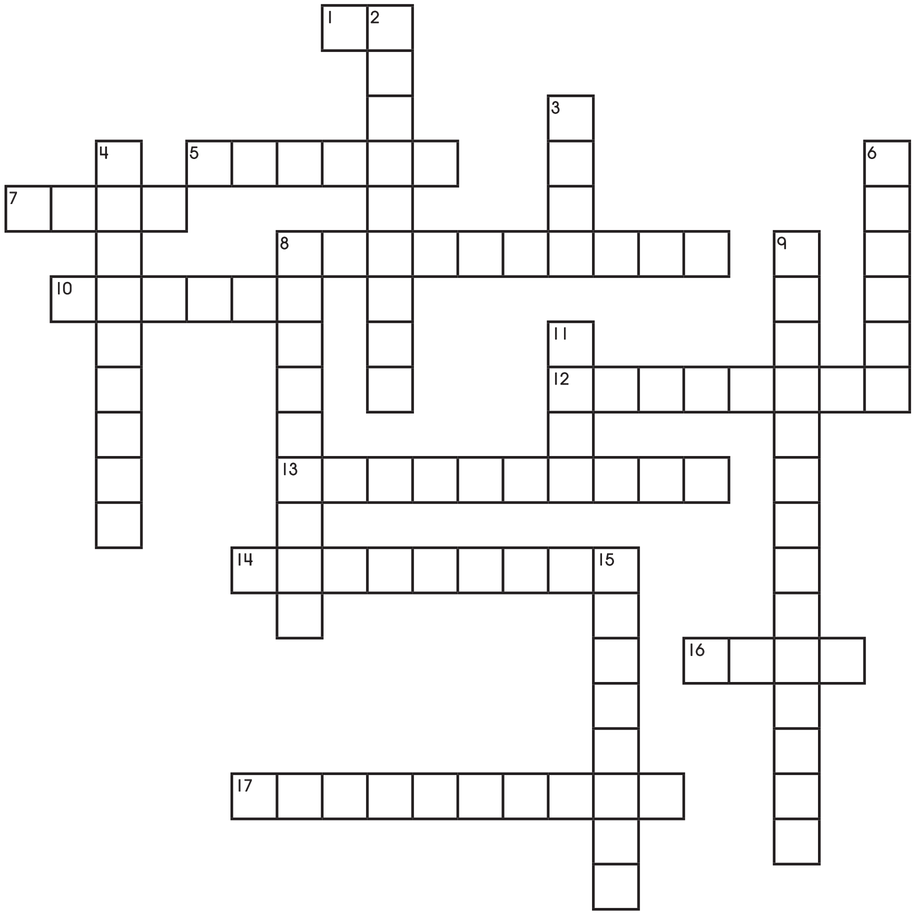
A weak acid and a weak base will produce any type of solution depending on the relative strengths of the acid and base involved.

Complete the chart for each salt shown.

Salt	Parent Acid	Parent Base	Type of Solution
1. KCl			
2. NH_4NO_3			
3. Na_3PO_4			
4. $CaSO_4$			
5. $AlBr_3$			
6. CuI_2			
7. MgF_4			
8. $NaNO_3$			
9. $LiC_2H_3O_2$			
10. $ZnCl_2$			
11. $SrSO_4$			
12. $Ba_3(PO_4)_2$			

Name________________________________

Acids and Bases Crossword



Across

1. Scale of acidity
5. An acid that consists of only two elements
7. Substance that forms hydronium ions in water (Arrhenius ____________)
8. This happens when an acid dissolves in water.
10. According to Bronsted-Lowry, an acid is a ________________ donor.
12. According to Bronsted-Lowry, a base is a proton ________________.
13. Can act as either an acid or a base
14. These pairs differ only by a proton.
16. An acid with a small K_a value would be a ________________ acid.
17. Reaction of an ion with H_2O to produce $H^+(aq) + OH^-(aq)$

Down

2. H_3O^+
3. Formed from the reaction of an acid and a base
4. Procedure to determine the concentration of an acid or base
6. A solution that will resist changes in pH
8. Changes color at the endpoint of a titration
9. The reaction of an acid with a base
11. Substance that produces hydroxide ions in aqueous solution (Arrhenius ____________)
15. When equivalent amounts of H^+ and OH^- have reacted in a titration

Name________________________________

Assigning Oxidation Numbers

Assign oxidation numbers to all of the elements in each compound or ion shown.

1. HCl	11. H_2SO_3
2. KNO_3	12. H_2SO_4
3. OH^-	13. BaO_2
4. Mg_3N_2	14. $KMnO_4$
5. $KClO_3$	15. LiH
6. $Al(NO_3)_3$	16. MnO_2
7. S_8	17. OF_2
8. H_2O_2	18. SO_3
9. PbO_2	19. NH_3
10. $NaHSO_4$	20. Na

Name______________________________

Redox Reactions

For each equation, identify the substance oxidized, the substance reduced, the oxidizing agent, and the reducing agent. Then, write the oxidation and reduction half-reactions.

Example:

oxidized		reduced		
Mg	+	Br_2	$\rightarrow$	$MgBr_2$
reducing agent		oxidizing agent		

oxidation half-reaction: $Mg^{\circ} \rightarrow Mg^{+2} + 2e^-$

reduction half-reaction: $2e^- + Br_2^{\circ} \rightarrow 2Br^-$

1. $2H_2 + O_2 \rightarrow 2H_2O$

2. $Fe + Zn^{2+} \rightarrow Fe^{2+} + Zn$

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

4. $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$

Name________________________________

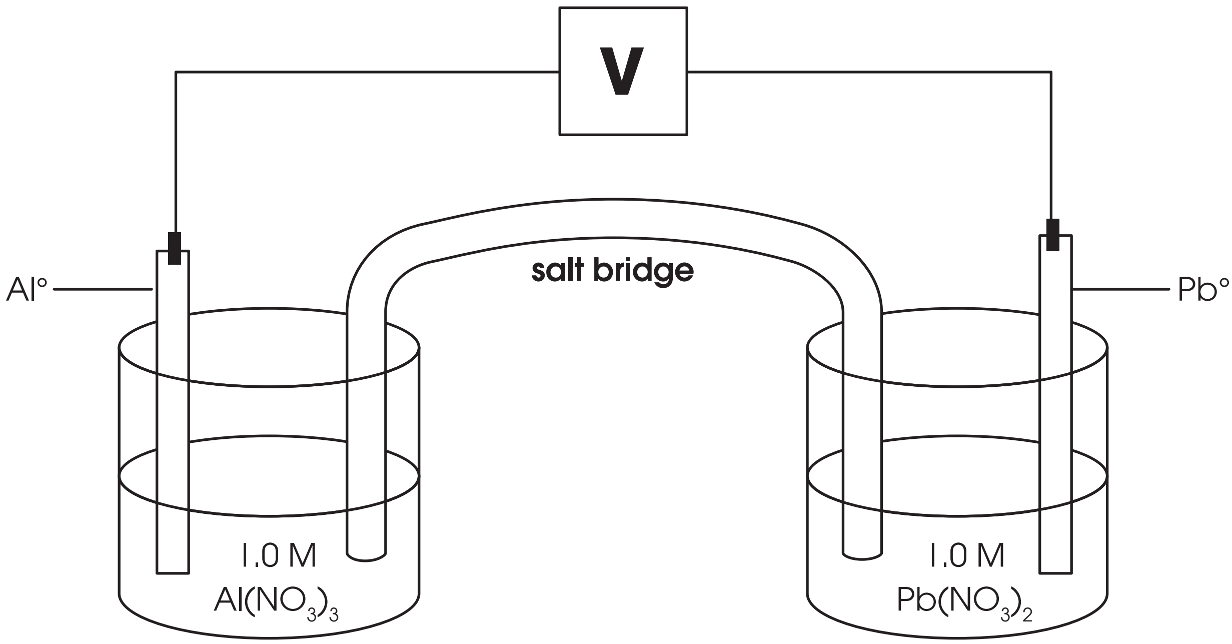
Balancing Redox Equations

Balance each equation using the half-reaction method.

1.	$Sn^{\circ} + Ag^{+} \rightarrow Sn^{2+} + Ag^{\circ}$
2.	$Cr^{\circ} + Pb^{2+} \rightarrow Cr^{3+} + Pb^{\circ}$
3.	$KClO_3 \rightarrow KCl + O_2$
4.	$NH_3 + O_2 \rightarrow NO + H_2O$
5.	$PbS + H_2O_2 \rightarrow PbSO_4 + H_2O$
6.	$H_2S + HNO_3 \rightarrow S + NO + H_2O$
7.	$MnO_2 + H_2C_2O_4 + H_2SO_4 \rightarrow MnSO_4 + CO_2 + H_2O$
8.	$H_2S + H_2SO_3 \rightarrow S + H_2O$
9.	$KIO_3 + H_2SO_3 \rightarrow KI + H_2SO_4$
10.	$K_2Cr_2O_7 + HCl \rightarrow KCl + CrCl_3 + Cl_2 + H_2O$

Name___________________________________

The Electrochemical Cell



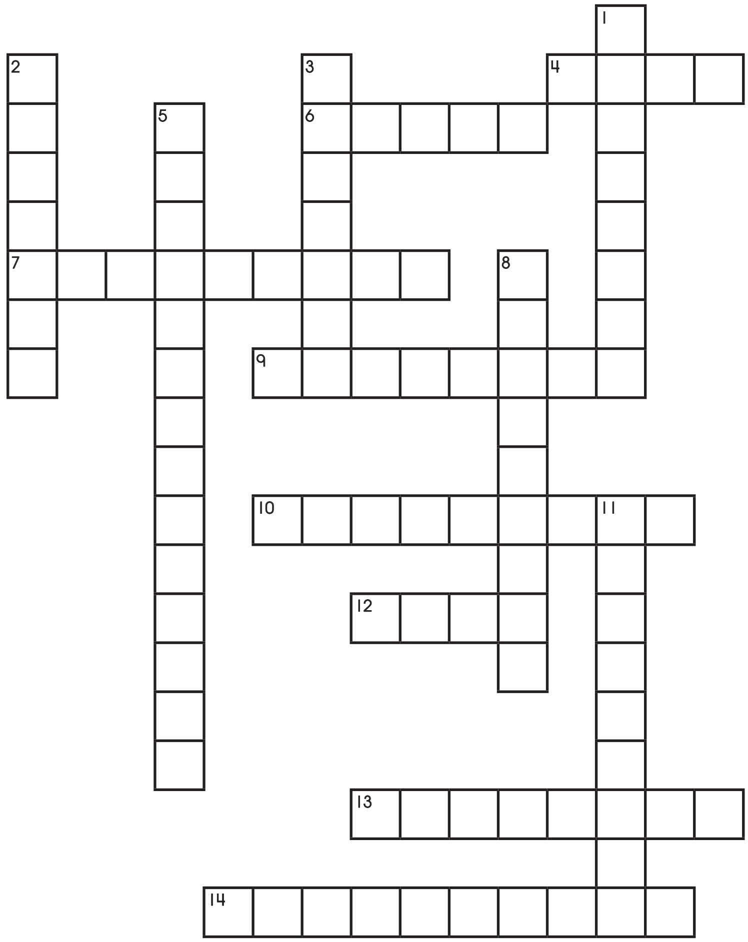
Answer each question, referring to the diagram and a table of standard electrode potentials.

1. Which is the more easily oxidized metal: aluminum or lead? ___________________
2. What is the balanced equation showing the spontaneous reaction that occurs?

   ____
3. What is the maximum voltage that the above cell can produce? ___________________
4. What is the direction of electron flow in the wire? ___________________
5. What is the direction of positive ion flow in the salt bridge? ___________________
6. Which electrode is decreasing in size? ___________________
7. Which electrode is increasing in size? ___________________
8. What is happening to the concentration of aluminum ions? ___________________
9. What is happening to the concentration of lead ions? ___________________
10. What is the voltage in this cell when the reaction reaches equilibrium? ___________________
11. Which is the anode? ___________________
12. Which is the cathode? ___________________
13. Which is the positive electrode? ___________________
14. Which is the negative electrode? ___________________

Name________________________________

Electrochemistry Crossword



ACROSS

4. Unit of electrical potential
6. Electrode where oxidation tokes place
7. Both atoms and ________ must be balanced in a redox equation.
9. The anode in an electrochemical cell has this charge.
10. Gain of electrons
12. Voltage of an electrochemical cell when it reaches equilibrium
13. A substance that is oxidized is the ________ agent.
14. Allows the flow of ions in an electrochemical cell

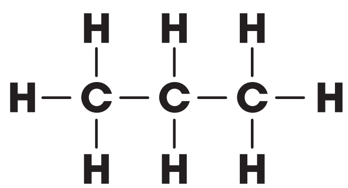
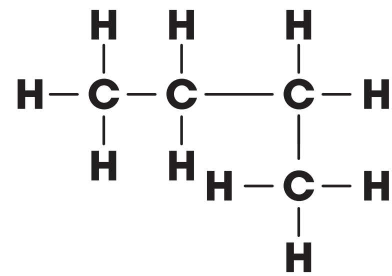
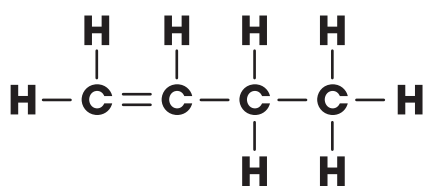
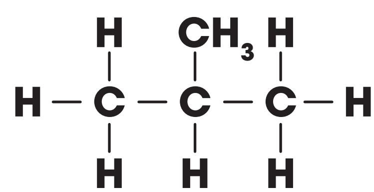
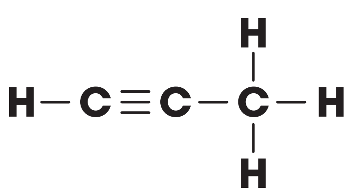
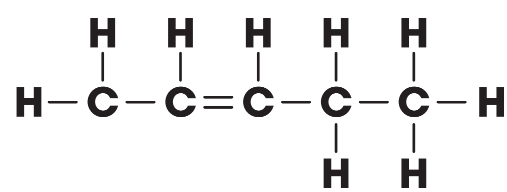
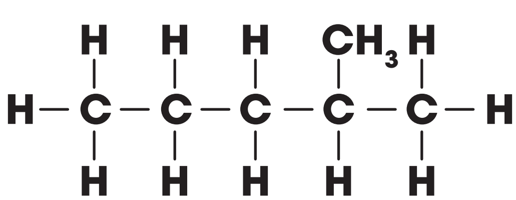
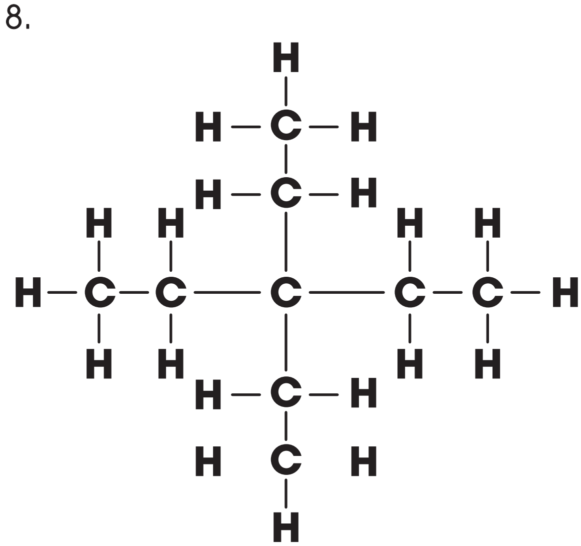
DOWN

1. The anode in an electrolytic cell has this charge.
2. Another word for on electrochemical cell
3. Electrode where reduction takes place
5. Process of layering a metal onto a surface in an electrolytic cell
8. Loss of electrons
11. A substance that is reduced is the ________ agent.

Name________________________________

Naming Hydrocarbons

Name each compound according to the IUPAC naming system.

1. 	5. 
2. 	6. 
3. 	7. 
4. 	8. 

Name________________________________

Structure of Hydrocarbons

Draw the structure of each compound.

1. ethane	5. ethyne
2. propene	6. 3, 3-dimethylpentane
3. 2-butene	7. 2, 3-dimethylpentane
4. methane	8. n-butyne

Name________________________________

Functional Groups

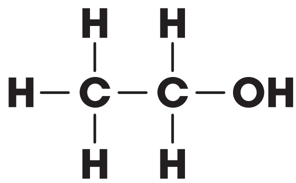
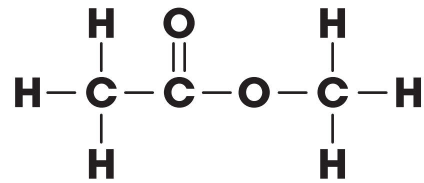
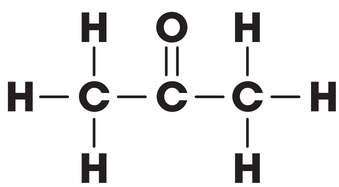
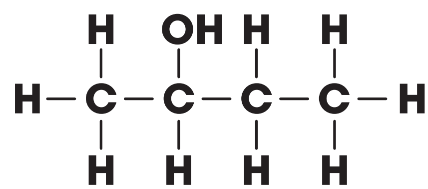
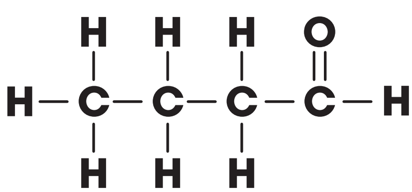
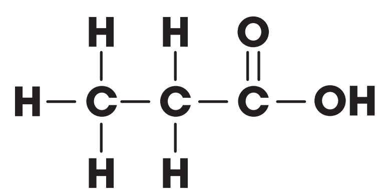
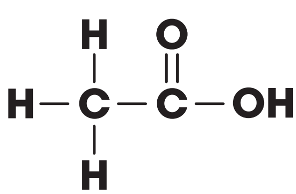
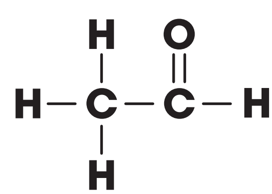
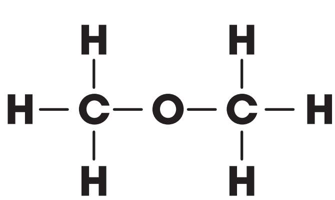
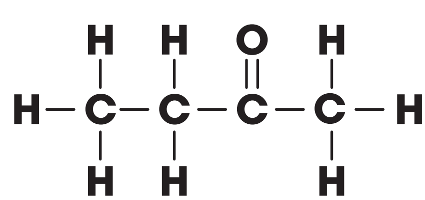
Classify each of the organic compounds as an *alcohol, carboxylic acid, aldehyde, ketone, ether*, or *ester*. Then, draw its structural formula.

1. CH_3COOH	6. $CH_3CH(OH)CH_3$
2. CH_3COCH_3	7. CH_3CH_2COOH
3. CH_3CH_2OH	8. $CH_3CH_2COOCH_3$
4. $CH_3CH_2OCH_3$	9. $CH_3CH_2COCH_3$
5. CH_3CH_2CHO	10. CH_3OCH_3

Name________________________________

Naming Other Organic Compounds

Name each compound.

1. 	6. 
2. 	7. 
3. 	8. 
4. 	9. 
5. 	10. 

Name_______________________________

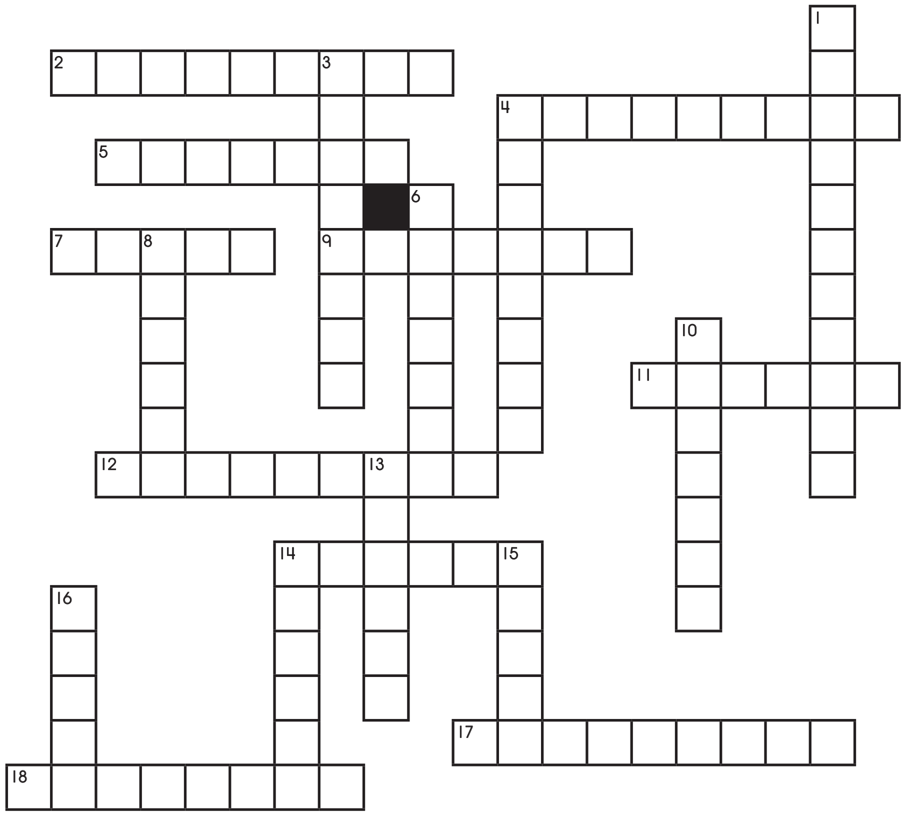
Structures of Other Organic Compounds

Draw the structure of each compound.

1. butanoic acid	6. methyl methanoate (methyl formate)
2. methanal	7. 3-pentanol
3. methanol	8. methanoic acid (formic acid)
4. butanone	9. propanal
5. diethyl ether	10. 2-pentanone

Name________________________________

Organic Chemistry Crossword



Across

2. Hydrocarbons containing only single bonds
4. Contains two double bonds
5. Alcohol in which the hydroxyl group is attached to an end carbon
7. An alkane minus one hydrogen; It attaches to another carbon chain.
9. Compounds with the same molecular formula, but different structural formulas
11. A dihydroxy alcohol
12. Alcohol in which the hydroxyl is attached to a carbon attached to two other carbons
14. Open chain hydrocarbon containing one double bond
17. Organic compounds containing the benzene ring structure
18. Describes a hydrocarbon with a side chain of carbon atoms

Down

1. An alcohol with only one hydroxyl group in its structure
3. Alcohol in which the hydroxyl is attached to a carbon attached to three other carbons
4. General formula R-CHO
6. High molecular mass compound consisting of repeating units called monomers
8. General formula R-CO-R'
10. Contains one or more -OH groups
13. Saturated open chain hydrocarbon
14. Open chain hydrocarbon containing only one triple bond
15. Produced by the reaction of an alcohol and an acid
16. General formula R-O-R'

Name________________________________

Reference

Selected Polyatomic Ions

Hg_2^{2+}	dimercury(I)	CrO_4^{2-}	chromate
NH_4^+	ammonium	$Cr_2O_7^{2-}$	dichromate
$C_2H_3O_2^-$, CH_3COO^-	acetate	MnO_4^-	permanganate
		MnO_4^{2-}	manganate
CN^-	cyanide	NO_2^-	nitrite
CO_3^{2-}	carbonate	NO_3^-	nitrate
HCO_3^-	hydrogen carbonate	OH^-	hydroxide
$C_2O_4^{2-}$	oxalate	PO_4^{3-}	phosphate
ClO^-	hypochlorite	SCN	thiocyanate
ClO_2^-	chlorite	SO_3^{2-}	sulfite
ClO_3^-	chlorate	SO_4^{2-}	sulfate
ClO_4^-	perchlorate	HSO_4^-	hydrogen sulfate
		$S_2O_3^{2-}$	thiosulfate

Fundamental Concepts

Density $D = \frac{m}{V}$

Force $F = ma$

Pressure $P = \frac{F}{A}$

Temperature Conversion $°C = \frac{5}{9}(°F - 32)$ $°F = \left(\frac{5}{9}\right)°C + 32$

Work $w = Fd$

Solutions

Boiling Point Elevation $\Delta T_b = K_b m$

Dilution $V_1C_1 = V_2C_2$

Freezing Point Depression $\Delta T_f = \Delta T_f = K_f m$

Henry's Law $C_g = k_g P_g$ or $S = k_H P$

Molality $m = \frac{\text{mol solute}}{\text{kg solvent}}$

Molarity $M = \frac{\text{mol solute}}{\text{L solution}}$

Normality $N = \frac{\text{equivilants solute}}{\text{L solution}}$

Raoult's Law $P_A = P°_A X_A$ $\Delta P_A = P°_A X_B$

Gas Laws

Boyle's Law $P_1V_1 = P_2V_2$

Charles' Law $V_1T_2 = V_2T_1$

Combined Gas Law $V_1P_1T_2 = V_2P_2T_1$

Dalton's Law $P_T = P_1 + P_2 + P_3 \ldots$

Graham's Law $\frac{r_1}{r_2} = \frac{u_1}{u_2} = \sqrt{\frac{3RT/M_1}{3RT/M_2}} = \sqrt{\frac{M_2}{M_1}}$

Ideal Gas Law $PV = nRT$

Van der Waals Equation $\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$

Energy Changes in Chemical Reactions

Enthalpy $\Delta H = H_{products} - H_{reactants}$

$\Delta E = q + w = q - P\Delta V$

$qp = \Delta E + P\Delta V = \Delta H$

Entropy $\Delta S = S_{products} - S_{reactants}$

Free Energy $\Delta G° = \Delta H° - T(\Delta S°)$

Chemical Equilibrium

Equilibrium Constant Expression

for $aA + bB \rightleftharpoons cC + dD$

$$K_c = \frac{(C)^c(D)^d}{(A)^a(B)^b} = \frac{k_{forward\ reaction}}{k_{reverse\ reaction}}$$

Electrochemistry

Electromotive Force $E° = \left(\frac{2.303RT}{n\mathcal{F}}\right)\log K$

Nernst Equation $E = E° - \left(\frac{0.0591}{n}\right)\log Q$

Acidity and Basicity

Ion Product Constant of Water $k_w = (H^+)(OH^-) = 1.0 \times 10^{-14}$

pH $pH = -\log(H^+)$ $pOH = -\log(OH^-)$

$pH + pOH = -\log K_W = 14.00$

$K_aK_b = (H^+)(OH^-) = k_W$

Radiant Energy

Wave Velocity $c = \nu\lambda$

Energy of a Photon $E = h\nu$

Ionic Equilibrium

Henderson-Hasselbach Equation $pH = pK_a + \log\frac{(base)}{(acid)}$

Solubility Product Constant $K_{sp} = (A^+)^m(X^-)^n$

Behavior of Nuclear Particles

Mass-Energy Relationship $E = mc^2$

Wavelength of a Particle $\lambda = \frac{h}{m\nu}$

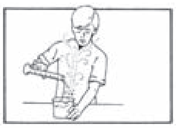
Answer Key

Name__________________________

Laboratory Dos And Don'ts

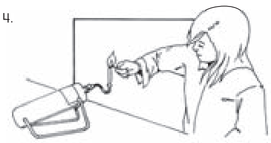
Identify what is wrong in each laboratory activity.

1.



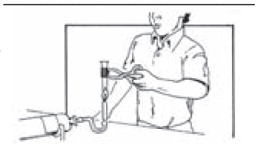
no goggles and no apron

4.



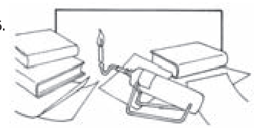
no goggles, dangling sleeves, no apron, hair not pulled back

2.



no goggles, no apron, heating from bottom, and perpendicular to desk

5.



flammable items too close to burner

3.

no goggles, no apron, sniffing directly from the test tube

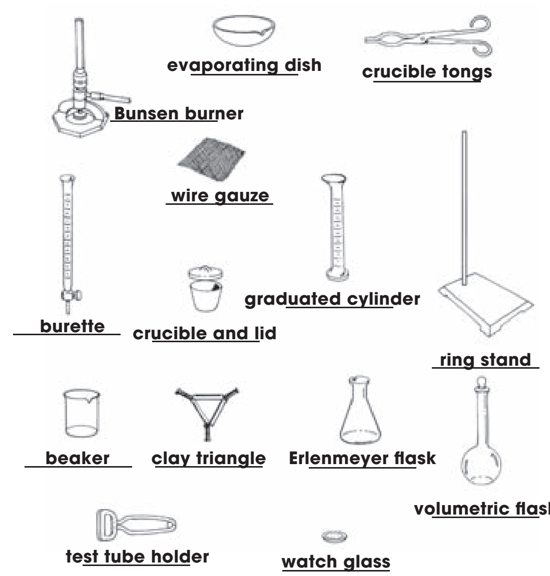
6.

no goggles, no apron, engaging in horseplay in the lab

Name__________________________

Laboratory Equipment

Label the lab equipment.

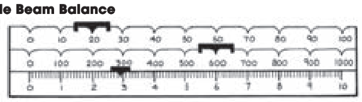


Name__________________________

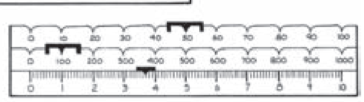
Triple And Four Beam Balances

Identify the mass on each balance.

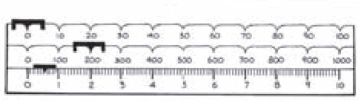
Triple Beam Balance



622.89 g

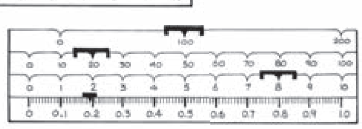


153.70 g

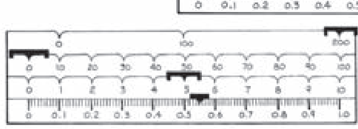


200.54 g

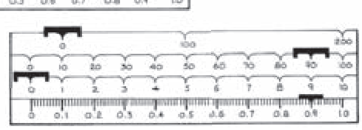
Four Beam Balance



128.192 g



205.548 g

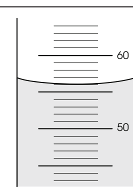


90.900 g

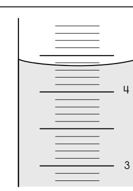
Name__________________________

Measuring Liquid Volume

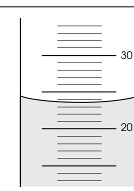
Identify the volume indicated on each graduated cylinder. The unit of volume is mL.



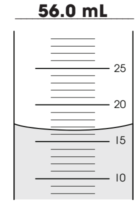
1. 56.0 mL



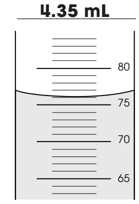
4. 4.35 mL



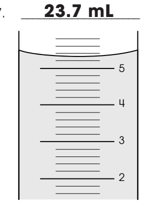
7. 23.7 mL



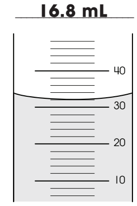
2. 16.8 mL



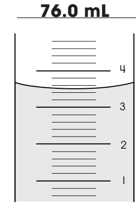
5. 76.0 mL



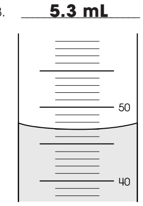
8. 5.3 mL



3. 31.0 mL



6. 3.5 mL



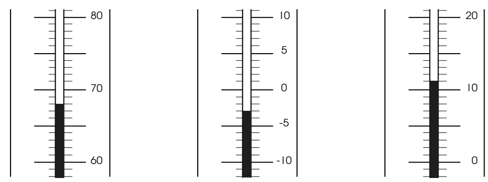
9. 47.0 mL

Answer Key

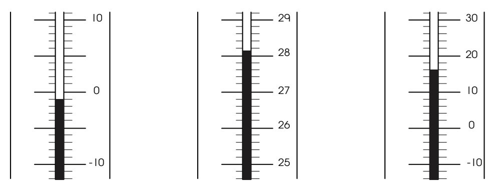
Name________________________

Reading Thermometers

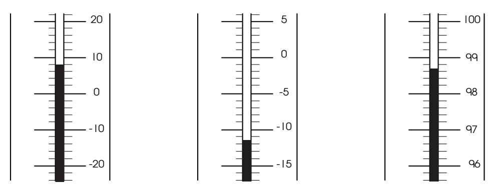
Identify the temperature indicated on each thermometer.



1. 68.0°
4. –3°
7. 11.0°



2. –1.1°
5. 28.1°
8. 16.0°



3. 8.0°
6. –11.5°
9. 98.7°

Name________________________

Metrics And Measurements

In the chemistry classroom and lab, the metric system of measurement is used. It is important to be able to convert from one unit to another.

mega-	kilo-	hecto-	deca-	Basic Units	deci-	centi-	milli-	micro-
(M)	(k)	(h)	(da)	gram (g)	(d)	(c)	(m)	(μ)
1,000,000	1,000	100	10	liter (L)	0.1	0.01	0.001	0.000001
10^6	10^3	10^2	10^1	meter (m)	10^{-1}	10^{-2}	10^{-3}	10^{-6}

Unit Factor Method

1. Write the given number and unit.
2. Set up a conversion factor (fraction used to convert one unit to another).
   a. Place the given unit as the denominator of the conversion factor.
   b. Place the desired unit as the numerator.
   c. Place a one in front of the larger unit.
   d. Determine the number of smaller units needed to make one of the larger units.
3. Cancel the units. Solve the problem.

Example 1: 55 mm = ________ m

$\frac{55\ \cancel{mm}}{} \cdot \frac{1\ m}{1{,}000\ \cancel{mm}} = 0.055\ m$

Example 2: 88 km = ________ m

$\frac{88\ \cancel{km}}{} \cdot \frac{1{,}000\ m}{1\ \cancel{km}} = 88{,}000\ m$

Example 3: 7,000 cm = ________ hm

$\frac{7{,}000\ \cancel{cm}}{} \cdot \frac{1\ \cancel{m}}{100\ \cancel{cm}} \cdot \frac{1\ hm}{10{,}000\ \cancel{cm}} = 0.7\ hm$

Example 4: 8 daL = ________ dL

$\frac{8\ \cancel{daL}}{} \cdot \frac{10\ \cancel{L}}{1\ \cancel{daL}} \cdot \frac{100\ dL}{1\ \cancel{daL}} = 800\ dL$

The unit factor method can be used to solve virtually any problem involving changes in units. It is especially useful in making complex conversions dealing with concentrations and derived units.

Convert each measurement.

1. 35 mL = 0.35 dL
2. 275 mm = 27.5 cm
3. 1,000 mL = 1.000 L
4. 25 cm = 250 mm
5. 0.075 m = 7.5 cm
6. 950 g = 0.950 kg
7. 1,000 L = 1.000 kL
8. 4,500 mg = 4.5 g
9. 0.005 kg = 0.5 dag
10. 15 g = 15,000 mg

Name________________________

Dimensional Analysis (Unit Factor Method)

Using this method, it is possible to solve many problems by using the relationship of one unit to another. For example, 12 inches = one foot. Since these two numbers represent the same value, the fractions 12 in./1 ft. and 1 ft./12 in. are both equal to one. When you multiply another number by the number one, you do not change its value. However, you may change its unit.

Example 1: Convert 2 miles to inches.

$2\ \text{miles} \times \frac{5{,}280\ \text{ft.}}{1\ \text{mile}} \times \frac{12\ \text{inches}}{1\ \text{ft.}} = 126{,}720\ \text{in.}$

Example 2: How many seconds are in 4 days?

$4\ \text{days} \times \frac{24\ \text{hrs.}}{1\ \text{day}} \times \frac{60\ \text{min.}}{1\ \text{hr.}} \times \frac{60\ \text{sec.}}{1\ \text{min.}} = 345{,}600\ \text{sec.}$

Solve each problem. Round irrational numbers to the thousandths place.

1. 3 hr. = 10,800 sec.
2. 0.035 mg = 0.0035 cg
3. 5.5 kg = 12.1 lb.
4. 2.5 yd. = 90 in.
5. 1.3 yr. = 28,470 hr.
6. 3 moles = 1.806×10^{24} molecules (1 mole = 6.02×10^{23} molecules)
7. 2.5×10^{24} molecules = 4.152 moles
8. 5 moles = 112 liters (1 mole = 22.4 liters)
9. 100. liters = 4.46 moles
10. 50. liters = 1.344×10^{24} molecules
11. 5.0×10^{24} molecules = 186.047 liters
12. 7.5×10^3 mL = 7.5 liters

Name________________________

Scientific Notation

Scientists very often deal with very small and very large numbers, which can lead to a lot of confusion when counting zeros. We can express these numbers as powers of 10.

Scientific notation takes the form of $M \times 10^n$ where $1 \leq M < 10$ and n represents the number of decimal places to be moved. Positive n indicates the standard form is a large number. Negative n indicates a number between zero and one.

Example 1: Convert 1,500,000 to scientific notation.

Move the decimal point so that there is only one digit to its left, for a total of 6 places.

$1{,}500{,}000 = 1.5 \times 10^6$

Example 2: Convert 0.000025 to scientific notation.

For this, move the decimal point 5 places to the right.

$0.000025 = 2.5 \times 10^{-5}$

(Note that when a number starts out less than one, the exponent is always negative.)

Convert each number to scientific notation.

1. 0.005 = 5×10^{-3}
2. 5,050 = 5.05×10^3
3. 0.0008 = 8×10^{-4}
4. 1,000 = 1×10^3
5. 1,000,000 = 1×10^6
6. 0.25 = 2.5×10^{-1}
7. 0.025 = 2.5×10^{-2}
8. 0.0025 = 2.5×10^{-3}
9. 500 = 5×10^2
10. 5,000 = 5×10^3

Convert each number to standard notation.

11. 1.5×10^3 = 1,500
12. 1.5×10^{-3} = 0.0015
13. 3.75×10^{-2} = 0.0375
14. 3.75×10^2 = 375
15. 2.2×10^5 = 220,000
16. 3.35×10^{-1} = 0.335
17. 1.2×10^{-4} = 0.00012
18. 1×10^4 = 10,000
19. 1×10^{-1} = 0.1
20. 4×10^0 = 4

Answer Key

Name______________________________

Significant Figures

A measurement can only be as accurate and precise as the instrument that produced it. A scientist must be able to express the accuracy of a number, not just its numerical value. We can determine the accuracy of a number by the number of significant figures it contains.

1. All digits 1–9 inclusive are significant.
   Example: 129 has 3 significant figures.
2. Zeros between significant digits are always significant.
   Example: 5,007 has 4 significant figures.
3. Trailing zeros in a number are significant only if the number contains a decimal point. Sometimes, a decimal may be added without any number in the tenths place.
   Example: 100.0 has 4 significant figures.
   100. has 3 significant figures.
   100 has 1 significant figure.
4. Zeros in the beginning of a number whose only function is to place the decimal point are not significant.
   Example: 0.0025 has 2 significant figures.
5. Zeros following a decimal significant figure are significant.
   Example: 0.000470 has 3 significant figures.
   0.47000 has 5 significant figures.

Determine the number of significant figures in each number.

1. 0.02 **1**
2. 0.020 **2**
3. 501 **3**
4. 501.0 **4**
5. 5,000 **1**
6. 5,000. **4**
7. 6,051.00 **6**
8. 0.0005 **1**
9. 0.1020 **4**
10. 10,001 **5**

Determine the location of the last significant place value by placing a bar over the digit. (Example: $1.70\bar{0}$)

11. $8,0\bar{4}0$
12. $0.9010\bar{0}$
13. $3.01 \times 1\bar{0}^{21}$
14. $0.030\bar{0}$
15. $90,\bar{1}00$
16. $0.00041\bar{0}$
17. $699.\bar{5}$
18. $4.\bar{7} \times 10^{-8}$
19. $2.00\bar{0} \times 10^{2}$
20. $10,800,0\bar{0}0.0$

Name______________________________

Calculations Using Significant Figures

When multiplying and dividing, limit and round to the least number of significant figures in any of the factors.

Example 1: 23.0 cm × 432 cm × 19 cm = 188,784 cm^3

The answer is expressed as 190,000 cm^3 since 19 cm has only two significant figures.

When adding and subtracting, limit and round your answer to the least number of decimal places in any of the numbers that make up your answer.

Example 2: 123.25 mL + 46.0 mL + 86.257 mL = 255.507 mL

The answer is expressed as 255.5 mL since 46.0 mL has only one decimal place.

Perform each operation, expressing the answer in the correct number of significant figures.

1. 1.35 m × 2.467 m = **3.33 m^2**
2. 1,035 m^2 ÷ 42 m = **25 m**
3. 12.01 mL + 35.2 mL + 6 mL = **53 mL**
4. 55.46 g – 28.9 g = **26.6 g**
5. 0.021 cm × 3.2 cm × 100.1 cm = **6.7 cm^3**
6. 0.15 cm + 1.15 cm + 2.051 cm = **3.35 cm**
7. 150 L^3 ÷ 4 L = **40 L^2**
8. 505 kg – 450.25 kg = **55 kg**
9. 1.252 mm × 0.115 mm × 0.012 mm = **0.0017 mm^3**
10. 1.278×10^3 m^2 ÷ 1.4267×10^2 m = **8.958 m**

Name______________________________

Percentage Error

Percentage error is a way for scientists to express how far off a laboratory value is from the commonly accepted value.

The formula is:

$$\% \text{ error} = \left| \frac{\text{Accepted Value} - \text{Experimental Value}}{\text{Accepted Value}} \right| \times 100$$

absolute value

Determine the percentage error in each problem.

1. Experimental value = 1.24 g Accepted value = 1.30 g	**4.62%**
2. Experimental value = 1.24×10^{-2} g Accepted value = 9.98×10^{-3} g	**24.2%**
3. Experimental value = 252 mL Accepted value = 225 mL	**12.0 %**
4. Experimental value = 22.2 L Accepted value = 22.4 L	**0.893%**
5. Experimental value = 125.2 mg Accepted value = 124.8 mg	**0.3%**

Name______________________________

Temperature and Its Measurement

Temperature (which measures average kinetic energy of the molecules) can be measured using three common scales: **Celsius**, **Kelvin**, and **Fahrenheit**. Use the following formulas to convert from one scale to another. Celsius is the scale most desirable for laboratory work. Kelvin represents the absolute scale. Fahrenheit is the old English scale, which is rarely used in laboratories.

°C = K – 273 K = °C + 273

$°F = \frac{9}{5}°C + 32$ $°C = \frac{5}{9}(°F - 32)$

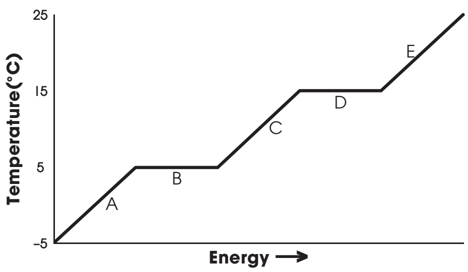
Complete the chart. All measurements are good to 1°C or better.

	°C	K	°F
1.	0°C	**273 K**	**32°F**
2.	**100°C**	**373 K**	212°F
3.	**177°C**	450 K	**351°F**
4.	**37.0°C**	**310 K**	98.6°F
5.	-273°C	**0 K**	**–459°F**
6.	**21°C**	294 K	**70°F**
7.	**25°C**	**298 K**	77°F
8.	**–48°C**	225 K	**–54°F**
9.	-40°C	**233 K**	**–40°F**

Answer Key

Name________________

Freezing and Boiling Point Graph



Use the graph to answer each question.

1. Which is the freezing point of the substance? **5°C**
1. Which is the boiling point of the substance? **15°C**
3. Which is the melting point of the substance? **5°C**
4. Which letter represents the range where the solid is being warmed? **A**
5. Which letter represents the range where the liquid is being warmed? **C**
6. Which letter represents the range where the vapor is being warmed? **E**
7. Which letter represents the melting of the solid? **B**
8. Which letter represents the vaporization of the liquid? **D**
9. Which letter(s) shows a change in potential energy? **B, D**
10. Which letter(s) shows a change in kinetic energy? **A, C, E**
11. Which letter represents condensation? **D**
12. Which letter represents crystallization? **B**

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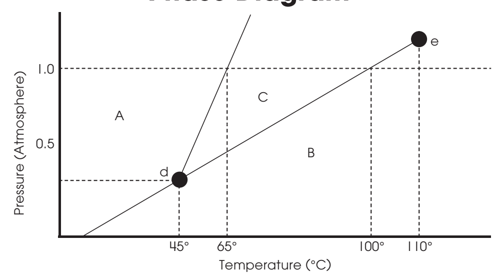
Name________________

Phase Diagram



Use the diagram to answer each question.

1. Which section represents the solid phase? **A**
2. Which section represents the liquid phase? **C**
3. Which section represents the gas phase? **B**
4. Which letter represents the triple point? **d**
5. Which letter represents the critical point? **e**
6. What is this substance's normal melting point? **65°C**
7. What is this substance's normal boiling point? **100°C**
8. Above what temperature is it impossible to liquify this substance no matter what the pressure? **110°C**
9. At what temperature and pressure do all three phases coexist? **45°C, 0.3 atm**
10. Is the density of the solid greater than or less than the density of the liquid? **greater than**
11. Would an increase in pressure cause this substance to freeze or melt? **freeze**

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Name________________

Heat and Its Measurement

Heat (or energy) can be measured in units of **calories** or **joules**. When there is a temperature change (ΔT), heat (Q) can be calculated using this formula:

Q = mass × ΔT × specific heat capacity
(ΔT = final temperature – initial temperature)

During a phase change, use this formula:

Q = mass × heat of fusion (or heat of vaporization)

Solve each problem.

1.	How many joules of heat are given off when 5.0 g of water cool from 75°C to 25°C? (Specific heat of water = 4.18 J/g°C) **1,000 J**
2.	How many calories are given off by the water in problem 1? (Specific heat of water = 1.0 cal/g°C) **250 cal**
3.	How many joules does it take to melt 35 g of ice at 0°C? (heat of fusion = 333 J/g) **12,000 J**
4.	How many calories are given off when 85 g of steam condense to liquid water? (heat of vaporization = 539.4 cal/g) **46,000 cal**
5.	How many joules of heat are necessary to raise the temperature of 25 g of water from 10°C to 60°C? **5,000 j**
6.	How many calories are given off when 50 g of water at 0°C freezes? (heat of fusion = 79.72 cal/g) **4,000 cal**

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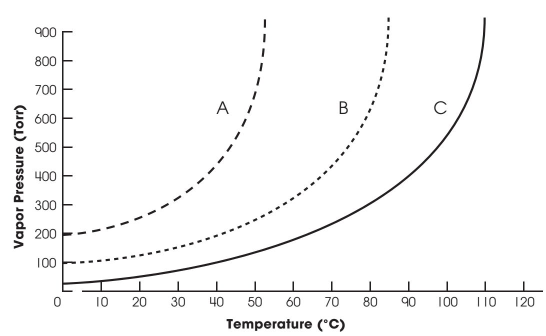
Name________________

Vapor Pressure and Boiling

A liquid will boil when its vapor pressure equals the atmospheric pressure.



Use the graph to answer each question.

1. At what temperature would Liquid A boil at an atmospheric pressure of 400 Torr? **40°C**
2. Liquid B? **70°C**
3. Liquid C? **92°C**
4. How low must the atmospheric pressure be for Liquid A to boil at 35°C? **375 Torr**
5. Liquid B? **150 Torr**
6. Liquid C? **75 Torr**
7. What is the normal boiling point of Liquid A? **50°C**
8. Liquid B? **82°C**
9. Liquid C? **108°C**
10. Which liquid has the strongest intermolecular forces? **C**

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Answer Key

Name________________________

Matter—Substances vs. Mixtures

All matter can be classified as either a **substance** (element or compound) or a **mixture** (heterogeneous or homogeneous).

Matter

- **Substance** — can write chemical formula, homogeneous
  - **Element** — one type of atom
  - **Compound** — two or more different atoms, chemically bonded
- **Mixture** — variable ratio
  - **Homogeneous** — solutions
  - **Heterogeneous** — colloids and suspensions

Classify each of the following as a substance or a mixture. If it is a substance, write *element* or *compound* in the substance column. If it is a mixture, write *heterogeneous* or *homogeneous* in the mixture column.

Type of Matter	Substance	Mixture
1. chlorine	**element**	
2. water	**compound**	
3. soil		**heterogeneous**
4. sugar water		**homogeneous**
5. oxygen	**element**	
6. carbon dioxide	**compound**	
7. rocky road ice cream		**heterogeneous**
8. alcohol	**compound**	
9. pure air		**homogeneous**
10. iron	**element**	

Name________________________

Physical vs. Chemical Properties

A **physical property** is observed with the senses and can be determined without destroying the object. Color, shape, mass, length, and odor are all examples of physical properties.

A **chemical property** indicates how a substance reacts with something else. The original substance is fundamentally changed in observing a chemical property. For example, the ability of iron to rust is a chemical property. The iron has reacted with oxygen, and the original iron metal is changed. It now exists as iron oxide, a different substance.

Classify each property as either chemical or physical by putting a check in the appropriate column.

	Physical Property	Chemical Property
1. blue color	✓	
2. density	✓	
3. flammability		✓
4. solubility	✓	
5. reacts with acid to form H_2		✓
6. supports combustion		✓
7. sour taste	✓	
8. melting point	✓	
9. reacts with water to form a gas		✓
10. reacts with a base to form water		✓
11. hardness	✓	
12. boiling point	✓	
13. can neutralize a base		✓
14. luster	✓	
15. odor	✓	

Name________________________

Physical vs. Chemical Changes

In a **physical change**, the original substance still exists; it only changes in form. In a **chemical change**, a new substance is produced. Energy changes always accompany chemical changes.

Classify each as a *physical* or *chemical* change.

1. Sodium hydroxide dissolves in water. **physical**
2. Hydrochloric acid reacts with potassium hydroxide to produce a salt, water, and heat. **chemical**
3. A pellet of sodium is sliced in two. **physical**
4. Water is heated and changed to steam. **physical**
5. Potassium chlorate decomposes to potassium chloride and oxygen gas. **chemical**
6. Iron rusts. **chemical**
7. When placed in H_2O, a sodium pellet catches on fire as hydrogen gas is liberated and sodium hydroxide forms. **chemical**
8. Water evaporates. **physical**
9. Ice melts. **physical**
10. Milk sours. **chemical**
11. Sugar dissolves in water. **physical**
12. Wood rots. **chemical**
13. Pancakes are cooking on a griddle. **chemical**
14. Grass is growing in a lawn. **chemical**
15. A tire is inflated with air. **physical**
16. Food is digested in the stomach. **chemical**
17. Water is absorbed by a paper towel. **physical**

Name________________________

Boyle's Law

Boyle's Law states that the volume of a given sample of gas at a constant temperature varies inversely with the pressure. (If one goes up, the other goes down.) Use the formula:

$$P_1 \times V_1 = P_2 \times V_2$$

Solve each problem (assuming constant temperature).

1. A sample of oxygen gas occupies a volume of 250. mL at 740. Torr. What volume will it occupy at 800. Torr pressure? **$23\bar{1}$ mL**
2. A sample of carbon dioxide occupies a volume of 3.50 liters at 125 kPa pressure. What pressure would the gas exert if the volume was decreased to 2.00 liters? **$2\bar{1}9$ kPa**
3. A 2.0 liter container of nitrogen has a pressure of 3.2 atm. What volume would be necessary to decrease the pressure to 1.0 atm? **6.4 L**
4. Ammonia gas occupies a volume of 450. mL at a pressure of 720. mmHg. What volume will it occupy at standard pressure? **$42\bar{6}$ mL**
5. A 175 mL sample of neon has its pressure changed from 75 kPa to 150 kPa. What is its new volume? **$\bar{8}8$ mL**
6. A sample of hydrogen at 1.5 atm has its pressure decreased to 0.50 atm, producing a new volume of 750 mL. What was its original volume? **250 mL**
7. Chlorine gas occupies a volume of 1.2 liters at 720 Torr. What volume will it occupy at 1 atm pressure? **$1.\bar{1}$ L**
8. Fluorine gas exerts a pressure of 900. Torr. When the pressure is changed to 1.50 atm, its volume is 250. mL. What was the original volume? **$3\bar{1}7$ mL**

Answer Key

Name________________________

Charles' Law

Charles' Law states that the volume of a given sample of gas at a constant pressure is directly proportional to the temperature in Kelvin. Use the following formulas:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{or} \quad V_1 \times T_2 = V_2 \times T_1$$

$$K = {}^\circ C + 273$$

Solve each problem (assuming constant pressure).

1. A sample of nitrogen occupies a volume of 250 mL at 25°C. What volume will it occupy at 95°C?
 3$\bar{1}$0 mL
2. Oxygen gas is at a temperature of 40°C when it occupies a volume of 2.3 liters. To what temperature should it be raised to occupy a volume of 6.5 liters?
 8$\bar{8}$0 K or 6$\bar{1}$0°C
3. Hydrogen gas was cooled from 1$\bar{5}$0°C to $\bar{5}$0°C. Its new volume is 75 mL. What was its original volume?
 9$\bar{8}$ mL
4. Chlorine gas occupies a volume of 25 mL at 3$\bar{0}$0 K. What volume will it occupy at 600 K?
 50 mL
5. A sample of neon gas at $\bar{5}$0°C and a volume of 2.5 liters is cooled to 25°C. What is the new volume?
 2.$\bar{3}$ L
6. Fluorine gas at 3$\bar{0}$0 K occupies a volume of 5$\bar{0}$0 mL. To what temperature should it be lowered to bring the volume to 3$\bar{0}$0 mL?
 180 K or −93°C
7. Helium occupies a volume of 3.8 liters at -45°C. What volume will it occupy at 45°C?
 5.3 L
8. A sample of argon gas is cooled and its volume went from 3$\bar{8}$0 mL to 2$\bar{5}$0 mL. If its final temperature was -55°C, what was its original temperature?
 33$\bar{1}$ K or 58°C

Name________________________

Combined Gas Law

In practical terms, it is often difficult to hold any of the variables constant. When there is a change in pressure, volume, and temperature, the combined gas law is used.

$$\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2} \quad \text{or} \quad P_1V_1T_2 = P_2V_2T_1$$

Complete the chart.

	P_1	V_1	T_1	P_2	V_2	T_2
1.	1.5 atm	3.0 L	2$\bar{0}$°C	2.5 atm	**1.9 L**	3$\bar{0}$°C
2.	720 Torr	256 mL	25°C	**8$\bar{0}$0 Torr**	250 mL	5$\bar{0}$°C
3.	6$\bar{0}$0 mmHg	2.5 L	22°C	760 mmHg	1.8 L	**269 K or −4°C**
4.	**1.2 atm**	750 mL	0.0°C	2.0 atm	500 mL	25°C
5.	95 kPa	4.0 L	**295 K or 22°C**	101 kPa	6.0 L	471 K or 198°C
6.	650. Torr	**275 mL**	100°	900. Torr	225 mL	15$\bar{0}$°C
7.	850 mmHg	1.5 L	15°C	**540 mmHg**	2.5 L	3$\bar{0}$°C
8.	125 kPa	125 mL	**544 K or 271°C**	1$\bar{0}$0 kPa	1$\bar{0}$0 mL	75°C

Name________________________

Dalton's Law of Partial Pressures

Dalton's Law says that the sum of the individual pressures of all the gases that make up a mixture is equal to the total pressure, or: $P_T = P_1 + P_2 + P_3 + \ldots$ The partial pressure of each gas is equal to the mole fraction of each gas times the total pressure.

$$P_T = P_1 + P_2 + P_3 + \ldots \quad \text{or} \quad \frac{\text{moles gas}_x}{\text{total moles}} \times P_T = P_x$$

Solve each problem.

1. A 250. mL sample of oxygen is collected over water at 25°C and 760.0 Torr. What is the pressure of the dry gas alone? (Vapor pressure of water at 25°C = 23.8 Torr)
 736 Torr
2. A 32.0 mL sample of hydrogen is collected over water at 2$\bar{0}$°C and 750.0 Torr. What is the pressure of the dry gas alone? (Vapor pressure of water at 2$\bar{0}$°C = 17.5 Torr)
 732.5 Torr
3. A 54.0 mL sample of oxygen is collected over water at 23°C and 770.0 Torr. What is the pressure of the dry gas alone? (Vapor pressure of water at 23°C = 21.1 Torr)
 748.9 Torr
4. A mixture of 2.00 moles of H_2, 3.00 moles of NH_3, 4.00 moles of CO_2, and 5.00 moles of N_2 exerts a total pressure of 80$\bar{0}$ Torr. What is the partial pressure of each gas?
 H_2 = 114.2 Torr **CO_2 = 228.6 Torr**
 NH_3 = 171.4 Torr **N_2 = 285.7 Torr**
5. The partial pressure of F_2 is 300 Torr in a mixture of gases where the total pressure is 1.00 atm. If there are $\bar{1}$.5 total moles in the mixture, how many moles of F_2 are present?
 263 moles

Name________________________

Ideal Gas Law

The **ideal gas law** describes the state of an ideal gas. While an ideal gas is hypothetical, the ideal gas law can be used to approximate the behavior of many gases under normal conditions. Use the formula:

$PV = nRT$ where P = pressure in atmospheres
V = volume in liters
n = number of moles of gas
R = Universal Gas Constant 0.0821 L•atm/mol•K
T = Kelvin temperature

Use the ideal gas law to solve each problem.

1. How many moles of oxygen will occupy a volume of 2.5 liters at 1.2 atm and 25°C? **0.12 moles**
2. What volume will 2.0 moles of nitrogen occupy at 720 Torr and 2$\bar{0}$°C? **51 L**
3. What pressure will be exerted by 25 g of CO_2 at a temperature of 25°C and a volume of 5$\bar{0}$0 mL? **28 atm**
4. At what temperature will 5.00 g of Cl_2 exert a pressure of 900. Torr at a volume of 7$\bar{5}$0 mL? **154 K or −11$\bar{9}$°C**
5. What is the density of NH_3 at 8$\bar{0}$0 Torr and 25°C? **0.73 g/L**
6. If the density of a gas is 1.2 g/L at 745. Torr and 2$\bar{0}$°C, what is its molecular mass? **29 g/mol**
7. How many moles of nitrogen gas will occupy a volume of 347 mL at 6680 Torr and 27°C? **0.124 moles**
8. What volume will 454 grams (1 lb.) of hydrogen occupy at 1.05 atm and 25°C? **10,494.6 L**
9. Find the number of grams of CO_2 that exert a pressure of 785 Torr at a volume of 32.5 L and a temperature of 32°C. **23 g**
10. An elemental gas has a mass of 10.3 g. If the volume is 58.4 L and the pressure is 758 Torr at a temperature of 2.5°C, what is the gas? **Br_2 (bromine)**

Answer Key

Name________________________

Graham's Law of Effusion

Graham's Law states that a gas will effuse at a rate that is inversely proportional to the square root of its molecular mass, *MM*.

$$\frac{rate_1}{rate_2} = \sqrt{\frac{MM_2}{MM_1}}$$

Solve each problem.

1. Under the same conditions of temperature and pressure, how many times faster will hydrogen effuse compared to carbon dioxide?

   4.7

2. If the carbon dioxide in problem 1 takes 32 seconds to effuse, how long will the hydrogen take?

   6.8 s

3. What is the relative rate of effusion of NH_3 compared to helium? Does NH_3 effuse faster or slower than helium?

   0.49, slower

4. If the helium in problem 3 takes 20 seconds to effuse, how long will NH_3 take?

   9.7 s

5. An unknown gas effuses 0.25 times as fast as helium. What is the molecular mass of the unkown gas?

   64 g/mol

Name________________________

Element Symbols

An element symbol can stand for one atom of the element or one mole of atoms of the element. (One mole = 6.02×10^{23} atoms of an element.)

Write the symbol for each element.

1. oxygen **O**
2. hydrogen **H**
3. chlorine **Cl**
4. mercury **Hg**
5. fluorine **F**
6. barium **Ba**
7. helium **He**
8. uranium **U**
9. radon **Rn**
10. sulfur **S**
11. plutonium **Pu**
12. calcium **Ca**
13. radium **Ra**
14. cobalt **Co**
15. zinc **Zn**
16. arsenic **As**
17. lead **Pb**
18. iron **Fe**

Write the name of the element that corresponds with each symbol.

19. Kr **krypton**
20. K **potassium**
21. C **carbon**
22. Ne **neon**
23. Si **silicon**
24. Zr **zirconium**
25. Sn **tin**
26. Pt **platinum**
27. Na **sodium**
28. Al **aluminum**
29. Cu **copper**
30. Ag **silver**
31. P **phosphorus**
32. Mn **manganese**
33. I **iodine**
34. Au **gold**
35. Mg **magnesium**
36. Ni **nickel**
37. Br **bromine**
38. Hg **mercury**

Name________________________

Atomic Structure

An atom is made up of protons and neutrons (both found in the nucleus) and electrons (in the surrounding electron cloud). The **atomic number** is equal to the number of protons. The **mass number** is equal to the number of protons plus neutrons.

In a neutral atom, the number of protons equals the number of electrons. The **charge** on an ion indicates an imbalance between protons and electrons. Too many electrons produce a negative charge. Too few electrons produce a positive charge.

This structure can be written as part of a chemical symbol.

mass number → $^{15}_{7}N^{3+}$ ← *charge*

atomic number →

7 protons
8 neutrons (15 – 7)
4 electrons

Complete the chart.

Element/ Ion	Atomic Number	Atomic Mass	Mass Number	Protons	Neutrons	Electrons
H	**1**	**1.0079**	**1**	**1**	**0**	**1**
H^+	**1**	**1.0079**	**1**	**1**	**0**	**0**
$^{12}_{6}C$	**6**	**12.011**	**12**	**6**	**6**	**6**
$^{7}_{3}Li^+$	**3**	**6.941**	**7**	**3**	**4**	**2**
$^{35}_{17}Cl^-$	**17**	**35.453**	**35**	**17**	**18**	**18**
$^{39}_{19}K$	**19**	**39.0983**	**39**	**19**	**20**	**19**
$^{24}_{12}Mg^{2+}$	**12**	**24.305**	**24**	**12**	**12**	**10**
As^{3-}	**33**	**74.9216**	**75**	**33**	**42**	**36**
Ag	**47**	**107.868**	**108**	**47**	**61**	**47**
Ag^{1+}	**47**	**107.868**	**108**	**47**	**61**	**46**
S^{2-}	**16**	**32.06**	**32**	**16**	**16**	**18**
U	**92**	**238.029**	**238**	**92**	**146**	**92**

Name________________________

Isotopes and Average Atomic Mass

Elements come in a variety of **isotopes**, meaning they are made up of atoms with the same atomic number but different atomic masses. These atoms differ in the number of neutrons.

The **average atomic mass** is the weighted average of all of the isotopes of an element.

Example: A sample of cesium is 75% ^{133}Cs, 20% ^{132}Cs, and 5% ^{134}Cs. What is its average atomic mass?

Answer: $0.75 \times 133 = 99.75$
$0.20 \times 132 = 26.4$
$0.05 \times 134 = 6.7$
Total = 132.85 amu = average atomic mass

Determine the average atomic mass of each mixture of isotopes.

1. 80% ^{127}I, 17% ^{126}I, 3% ^{128}I

   139.7 amu

2. 50% ^{197}Au, 50% ^{198}Au

   197.5 amu

3. 15% ^{55}Fe, 85% ^{56}Fe

   55.85 amu

4. 99% ^{1}H, 0.8% ^{2}H, 0.2% ^{3}H

   1.012 amu

5. 95% ^{14}N, 3% ^{15}N, 2% ^{16}N

   14.07 amu

6. 98% ^{12}C, 2% ^{14}C

   12.04 amu

Answer Key

Name________________________

Electron Configuration (Level One)

Electrons are distributed in the electron cloud into principal energy levels (1, 2, 3, ...), sublevels (*s, p, d, f*), orbitals (*s* has 1, *p* has 3, *d* has 5, *f* has 7), and spin (two electrons allowed per orbital).

Example: Draw the electron configuration of sodium (atomic number 11).

$1s^2$ ↑↓ $2s^2$ ↑↓ $2p^6$ ↑↓ ↑↓ ↑↓ $3s^1$ ↑

Draw the electron configuration of each atom.

1. Cl
$1s^2$ ↑↓ $2s^2$ ↑↓ $2p^6$ ↑↓ ↑↓ ↑↓ $3s^2$ ↑↓ $3p^5$ ↑↓ ↑↓ ↑

2. N
$1s^2$ ↑↓ $2s^2$ ↑↓ $2p^3$ ↑ ↑ ↑

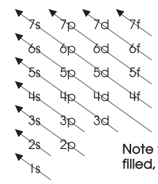
3. Al
$1s^2$ ↑↓ $2s^2$ ↑↓ $2p^6$ ↑↓ ↑↓ ↑↓ $3s^2$ ↑↓ $3p^1$ ↑ _ _

4. O
$1s^2$ ↑↓ $2s^2$ ↑↓ $2p^4$ ↑↓ ↑ ↑

Name________________________

Electron Configuration (Level Two)

At atomic numbers greater than 18, the sublevels begin to fill out of order. A good approximation of the order of filling can be determined using the diagonal rule.



Note that after the 3p sublevel is filled, the 4s is filled, and then the 3d.

Draw the electron configuration of each atom.

1. K
$1s^2$ ↑↓ $2s^2$ ↑↓ $2p^6$ ↑↓ ↑↓ ↑↓ $3s^2$ ↑↓ ↑↓ ↑↓ $3p^6$ ↑↓ ↑↓ ↑↓ $3d^0$ _ _ _ _ _ $4s^1$ ↑

2. V
$1s^2$ ↑↓ $2s^2$ ↑↓ $2p^6$ ↑↓ ↑↓ ↑↓ $3s^2$ ↑↓ $3p^6$ ↑↓ ↑↓ ↑↓ $3d^3$ ↑ ↑ ↑ $4s^2$ ↑↓

3. Co
$1s^2$ ↑↓ $2s^2$ ↑↓ $2p^6$ ↑↓ ↑↓ ↑↓ $3s^2$ ↑↓ $3p^6$ ↑↓ ↑↓ ↑↓ $3d^7$ ↑↓ ↑↓ ↑↓ ↑ ↑ $4s^2$ ↑↓

4. Zr
$1s^2$ ↑↓ $2s^2$ ↑↓ $2p^6$ ↑↓↑↓↑↓ $3s^2$ ↑↓ $3p^6$ ↑↓↑↓↑↓ $3d^{10}$ ↑↓↑↓↑↓↑↓↑↓ $4s^2$ ↑↓ $4p^6$ ↑↓↑↓↑↓ $4d^2$ ↑ ↑ _ _ _ $5s^2$ ↑↓

Name________________________

Valence Electrons

The **valence electrons** are the electrons in the outermost principal energy level. They are always *s* electrons or *s* and *p* electrons. Since the total number of electrons possible in *s* and *p* sublevels is eight, there can be no more than eight valence electrons.

Example: carbon

Electron configuration is $1s^2$ $\boxed{2s^2\ 2p^2}$.

Carbon has 4 valence electrons.

Determine the number of valence electrons in each atom.

1. fluorine ___7___
2. phosphorus ___5___
3. calcium ___2___
4. nitrogen ___5___
5. iron ___2___
6. argon ___8___
7. potassium ___1___
8. helium ___2___
9. magnesium ___2___
10. sulfur ___6___
11. lithium ___1___
12. zinc ___2___
13. carbon ___4___
14. iodine ___7___
15. oxygen ___6___
16. barium ___2___
17. aluminum ___3___
18. hydrogen ___1___
19. xenon ___8___
20. copper ___1___

Name________________________

Lewis Dot Diagrams

Lewis dot diagrams are a way to indicate the number of valence electrons around an atom.

Examples:

Draw the Lewis dot diagram of each atom.

1. calcium

2. potassium

3. argon

4. aluminum

5. bromine

6. carbon

7. helium

8. oxygen

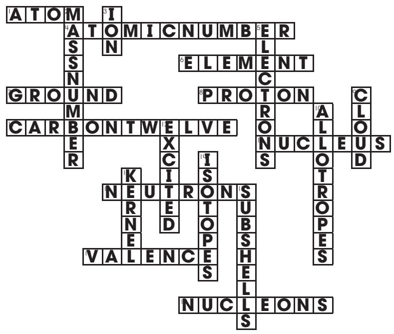
9. phosphorus

10. hydrogen

Answer Key

Name____________________

Atomic Structure Crossword



Across

1. The smallest particle of an element that can enter into chemical change
4. The number of protons in the nucleus of an atom
6. Cannot be decomposed into simpler substances by ordinary chemical means
7. State in which all electrons are at their lowest possible energy level
8. The positively charged particle found in the nucleus
11. Standard atomic mass unit for carbon
13. Most of the mass of an atom is here.
16. Mass number minus atomic number
18. Electrons in the outermost principal energy level
19. Protons and neutrons are these.

Down

2. Sum of the protons and neutrons in the nucleus of an atom
3. Charged atom or group of atoms
5. Equal to the number of protons in a neutral atom
9. The volume of an atom is determined by the size of its electron.
10. Different forms of the same element
12. State in which electrons have absorbed energy and "jumped" to a higher energy level
14. Atoms with the same atomic number but different atomic masses
15. The nucleus and all electrons in an atom except the valence electrons
17. *s, p, d, f*

Name____________________

Nuclear Decay

Predict the product of each nuclear reaction.

1. ${}^{42}K \longrightarrow {}^{0}_{-1}e +$ **${}^{42}_{20}Ca$**
2. ${}^{239}Pu \longrightarrow {}^{4}_{2}He +$ **${}^{235}_{92}U$**
3. ${}^{235}_{92}U \longrightarrow$ **${}^{4}_{2}He$** $+ {}^{231}_{90}Th$
4. ${}^{1}_{1}H + {}^{3}_{1}H \longrightarrow$ **${}^{4}_{2}He$**
5. ${}^{6}_{3}Li + {}^{1}_{0}n \longrightarrow {}^{4}_{2}He +$ **${}^{3}_{1}H$**
6. ${}^{27}_{13}Al + {}^{4}_{2}He \longrightarrow {}^{30}_{15}P +$ **${}^{1}_{0}N$**
7. ${}^{9}_{4}Be + {}^{1}_{1}H \longrightarrow$ **${}^{6}_{3}Li$** $+ {}^{4}_{2}He$
8. ${}^{37}K \longrightarrow {}^{0}_{+1}e +$ **${}^{37}_{18}Ar$**
9. **${}^{235}_{92}U$** $+ {}^{1}_{0}n \longrightarrow {}^{142}_{56}Ba + {}^{91}_{36}Kr + 3{}^{1}_{0}n$
10. ${}^{238}_{92}U + {}^{4}_{2}He \longrightarrow$ **${}^{241}_{94}Pu$** $+ {}^{1}_{0}n$

Name____________________

Half-Lives of Radioactive Isotopes

Solve each problem.

1.	How much of a 100.0 g sample of ${}^{198}Au$ is left after 8.10 days if its half-life is 2.70 days? **12.5 g**
2.	A 50.0 g sample of ${}^{16}N$ decays to 12.5 g in 14.4 seconds. What is its half-life? **7.2 s**
3.	The half-life of ${}^{42}K$ is 12.4 hours. How much of a 75$\bar{0}$ g sample is left after 62.0 hours? **23.4 g**
4.	What is the half-life of ${}^{99}Tc$ if a 50$\bar{0}$ g sample decays to 62.5 g in 639,000 years? **2.13×10^5 y**
5.	The half-life of ${}^{232}Th$ is 1.4×10^{10} years. If there are 25.0 g of the sample left after 2.8×10^{10} years, how many grams were in the original sample? **1$\bar{0}$0 g**
6.	There are 5.0 g of ${}^{131}I$ left after 40.35 days. How many grams were in the original sample if its half-life is 8.07 days? **160 g**

Name____________________

Periodic Table Worksheet

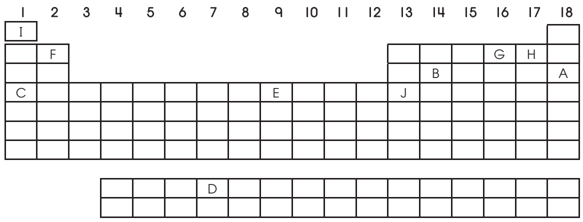
Use a copy of the periodic table to answer each question.

1. Where are the most active metals located? **lower left**
2. Where are the most active nonmetals located? **upper right**
3. As you go from left to right across a period, the atomic size (decreases, increases). Why? **increased positive nuclear charge**
4. As you travel down a group, the atomic size (decreases, increases). Why? **additional principal energy levels**
5. A negative ion is (larger, smaller) than its parent atom.
6. A positive ion is (larger, smaller) than its parent atom.
7. As you go from left to right across a period, the first ionization energy generally (decreases, increases). Why? **increased positive nuclear charge**
8. As you go down a group, the first ionization energy generally (decreases, increases). Why? **outermost electron is farther away from nucleus, shielding effect of inner electrons**
9. Where is the highest electronegativity found? **upper right (F)**
10. Where is the lowest electronegativity found? **lower left (Fr)**
11. Elements of Group 1 are called **alkali metals**.
12. Elements of Group 2 are called **alkaline earth metals**.
13. Elements of Group 3–12 are called **transition elements**.
14. As you go from left to right across the periodic table, the elements go from (metals, nonmetals) to (metals, nonmetals).
15. Group 17 elements are called **halogens**.
16. The most active element in Group 17 is **fluorine**.
17. Group 18 elements are called **noble gases**.
18. What sublevels are filling across the Transition Elements? ***d* and *f***
19. Elements within a group have a similar number of **valence electrons**.
20. Elements across a series have the same number of **principal energy levels**.
21. A colored ion generally indicates a **transition element**.
22. As you go down a group, the elements generally become (more, less) metallic.
23. The majority of elements in the periodic table are (metals, nonmetals).
24. Elements in the periodic table are arranged according to their **atomic numbers**
25. An element with both metallic and nonmetallic properties is called a **semimetal or metalloid**

Answer Key

Name____________________________

Periodic Table Puzzle

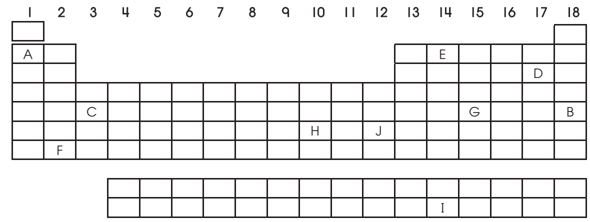


Place the letter of each of the above elements next to its description.

1. An alkali metal **C**
2. An alkaline earth metal **F**
3. An inactive gas **A**
4. An active nonmetal **H**
5. A semimetal **B**
6. An inner transition element **D**
7. Its most common oxidation state is -2. **G**
8. A metal with more than one oxidation state **E**
9. A metal with an oxidation number of +3 **J**
10. Has oxidation numbers of +1 and -1 **I**

Name____________________________

Periodic Table Puzzle



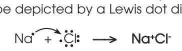
Place the letter of each of the above elements next to its description. Then, using a copy of the periodic table, write the element's name and symbol.

1. An element in Group 17 **D; chlorine, Cl**
2. An alkali metal with an oxidation of +1 **A; lithium, Li**
3. A metalloid **G; antimony, Sb**
4. Has 80 electrons **J; mercury, Hg**
5. Its electron configuration is 2, 4. **E; carbon, C**
6. Has 99 protons **I; einsteinium, Es**
7. An element in the fifth period with fewer than 50 protons **C; yttrium, Y**
8. Its electron configuration is 2, 8, 18, 18, 8. **B; xenon, Xe**
9. A transition metal with one electron in its outer shell **H; platinum, Pt**
10. An alkaline earth metal **F; radium, Ra**

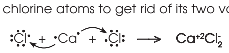
Name____________________________

Ionic Bonding

Ionic bonding occurs when a metal transfers one or more electrons to a nonmetal in an effort to attain a stable octet of electrons. For example, the transfer of an electron from sodium to chlorine can be depicted by a Lewis dot diagram.



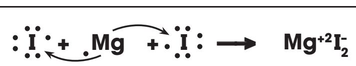
Calcium would need two chlorine atoms to get rid of its two valence electrons.



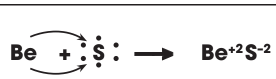
Sketch the transfer of electrons in each combination.

1. K + F

2. Mg + I

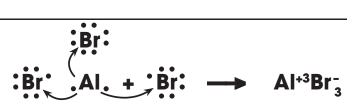


3. Be + S



4. Na + O

5. Al + Br

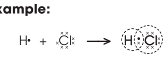


Name____________________________

Covalent Bonding

Covalent bonding occurs when two or more nonmetals share electrons, attempting to attain a stable octet of electrons at least part of the time.

Example:

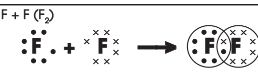


Note that hydrogen is content with 2, not 8, electrons.

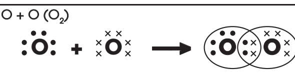
Sketch how covalent bonding occurs in each pair of atoms. Atoms may share one, two, or three pairs of electrons.

1. H + H (H_2)

2. F + F (F_2)

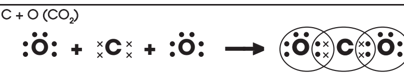


3. O + O (O_2)

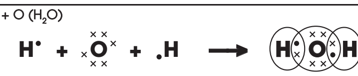


4. N + N (N_2)

5. C + O (CO_2)



6. H + O (H_2O)



Answer Key

Name____________________

Types of Chemical Bonds

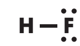
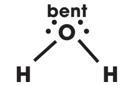
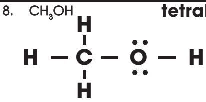
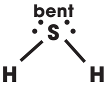
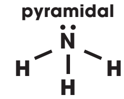
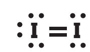
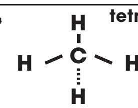
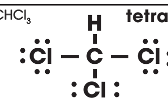
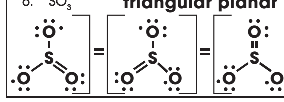
Identify each compound as *ionic* (metal + nonmetal), *covalent* (nonmetal + nonmetal), or *both* (compound containing a polyatomic ion).

1. $CaCl_2$ **ionic**
2. CO_2 **covalent**
3. H_2O **covalent**
4. $BaSO_4$ **both**
5. K_2O **ionic**
6. NaF **ionic**
7. Na_2CO_3 **both**
8. CH_4 **covalent**
9. SO_3 **covalent**
10. LiBr **ionic**
11. MgO **ionic**
12. NH_4Cl **both**
13. HCl **covalent**
14. KI **ionic**
15. NaOH **both**
16. NO_2 **covalent**
17. $AlPO_4$ **both**
18. $FeCl_3$ **ionic**
19. P_2O_5 **covalent**
20. N_2O_3 **covalent**

Name____________________

Shapes of Molecules

Using VSEPR theory, name and sketch the shape of each molecule.

1. N_2 **linear** 	7. HF **linear** 
2. H_2O **bent** 	8. CH_3OH **tetrahedral** 
3. CO_2 **linear** 	9. H_2S **bent** 
4. NH_3 **pyramidal** 	10. I_2 **linear** 
5. CH_4 **tetrahedral** 	11. $CHCl_3$ **tetrahedral** 
6. SO_3 **triangular planar** 	12. O_2 **linear** 

Name____________________

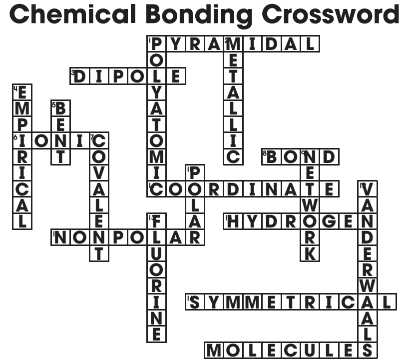
Polarity of Molecules

Identify each molecule as *polar* or *nonpolar*.

1. N_2 **nonpolar**	7. HF **polar**
2. H_2O **polar**	8. CH_3OH **polar**
3. CO_2 **nonpolar**	9. H_2S **polar**
4. NH_3 **polar**	10. I_2 **nonpolar**
5. CH_4 **nonpolar**	11. $CHCl_3$ **polar**
6. SO_3 **nonpolar**	12. O_2 **nonpolar**

Name____________________

Chemical Bonding Crossword



Across

1. Ammonia is polar because its shape is ________.
3. Word used to describe a molecule with an unequal charge distribution
6. Type of bond formed between an active metal and a nonmetal
8. The simultaneous attraction of electrons for the nuclei of two or more atoms is a chemical ________.
11. Type of covalent bond in which one atom donates both electrons
14. Bonding that is responsible for the relatively high boiling point of water
15. Type of covalent bond found in diatomic molecules
16. Carbon dioxide is nonpolar because it is ________.
17. Particles formed from covalent bonding

Down

1. Compounds with both ionic and covalent bonds contain this type of ion.
2. Type of bond found in aluminum foil
4. The formulas of ionic compounds must be expressed as ________ formulas.
5. The shape of a water molecule
7. Type of bond found between nonmetals
9. Type of covalent bonding that is found in a diamond
10. Type of covalent bond found between atoms of different electronegativity values
12. Force of attraction between nonpolar molecules
13. Element with the highest electronegativity value

Answer Key

Name________________________

Writing Formulas (Crisscross Method)

Write the formula of the compound produced from the listed ions.

	Cl^-	CO_3^{-2}	OH^-	SO_4^{-2}	PO_4^{-3}	NO_3^-
Na^+	$NaCl$	Na_2CO_3	$NaOH$	Na_2SO_4	Na_3PO_4	$NaNO_3$
NH_4^+	NH_4Cl	$(NH_4)_2CO_3$	NH_4OH	$(NH_4)_2SO_4$	$(NH_4)_3PO_4$	NH_4NO_3
K^+	KCl	K_2CO_3	KOH	K_2SO_4	K_3PO_4	KNO_3
Ca^{2+}	$CaCl_2$	$CaCO_3$	$Ca(OH)_2$	$CaSO_4$	$Ca_3(PO_4)_2$	$Ca(NO_3)_2$
Mg^{2+}	$MgCl_2$	$MgCO_3$	$Mg(OH)_2$	$MgSO_4$	$Mg_3(PO_4)_2$	$Mg(NO_3)_2$
Zn^{2+}	$ZnCl_2$	$ZnCO_3$	$Zn(OH)_2$	$ZnSO_4$	$Zn_3(PO_4)_2$	$Zn(NO_3)_2$
Fe^{3+}	$FeCl_3$	$Fe_2(CO_3)_3$	$Fe(OH)_3$	$Fe_2(SO_4)_3$	$FePO_4$	$Fe(NO_3)_3$
Al^{3+}	$AlCl_3$	$Al_2(CO_3)_3$	$Al(OH)_3$	$Al_2(SO_4)_3$	$AlPO_4$	$Al(NO_3)_3$
Co^{3+}	$CoCl_3$	$Co_2(CO_3)_3$	$Co(OH)_3$	$Co_2(SO_4)_3$	$CoPO_4$	$Co(NO_3)_3$
Fe^{2+}	$FeCl_2$	$FeCO_3$	$Fe(OH)_2$	$FeSO_4$	$Fe_3(PO_4)_2$	$Fe(NO_3)_2$
H^+	HCl	H_2CO_3	HOH or H_2O	H_2SO_4	H_3PO_4	HNO_3

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Name________________________

Naming Ionic Compounds

Name each compound using the Stock Naming System.

1. $CaCO_3$ calcium carbonate
2. KCl potassium chloride
3. $FeSO_4$ iron(II) sulfate
4. LiBr lithium bromide
5. $MgCl_2$ magnesium chloride
6. $FeCl_3$ iron(III) chloride
7. $Zn_3(PO_4)_2$ zinc phosphate
8. NH_4NO_3 ammonium nitrate
9. $Al(OH)_3$ aluminum hydroxide
10. $CuC_2H_3O_2$ copper(I) acetate
11. $PbSO_3$ lead(II) sulfite
12. $NaClO_3$ sodium chlorate
13. CaC_2O_4 calcium oxalate
14. Fe_2O_3 iron(III) oxide
15. $(NH_4)_3PO_4$ ammonium phosphate
16. $NaHSO_4$ sodium hydrogen sulfate or sodium bisulfate
17. Hg_2Cl_2 mercury(I) chloride
18. $Mg(NO_2)_2$ magnesium nitrate
19. $CuSO_4$ copper(II) sulfate
20. $NaHCO_3$ sodium hydrogen carbonate or sodium bicarbonate
21. $NiBr_3$ nickel(III) bromide
22. $Be(NO_3)_2$ beryllium nitrate
23. $ZnSO_4$ zinc sulfate
24. $AuCl_3$ gold(III) chloride
25. $KMnO_4$ potassium permanganate

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Name________________________

Naming Molecular Compounds

Name each covalent compound.

1. CO_2 carbon dioxide
2. CO carbon monoxide
3. SO_2 sulfur dioxide
4. SO_3 sulfur trioxide
5. N_2O dinitrogen monoxide
6. NO nitrogen monoxide
7. N_2O_3 dinitrogen trioxide
8. NO_2 nitrogen dioxide
9. N_2O_4 dinitrogen tetroxide
10. N_2O_5 dinitrogen pentoxide
11. PCl_3 phosphorus trichloride
12. PCl_5 phosphorus pentachloride
13. NH_3 ammonia
14. SCl_6 sulfur hexachloride
15. P_2O_5 diphosphorus pentoxide
16. CCl_4 carbon tetrachloride
17. SiO_2 silicon dioxide
18. CS_2 carbon disulfide
19. OF_2 oxygen difluoride
20. PBr_3 phosphorus tribromide

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Name________________________

Naming Acids

Name each acid.

1. HNO_3 nitric acid
2. HCl hydrochloric acid
3. H_2SO_4 sulfuric acid
4. H_2SO_3 sulfurous acid
5. $HC_2H_3O_2$ acetic acid
6. HBr hydrobromic acid
7. HNO_2 nitrous acid
8. H_3PO_4 phosphoric acid
9. H_2S hydrosulfuric acid
10. H_2CO_3 carbonic acid

Write the formula of each acid.

11. sulfuric acid H_2SO_4
12. nitric acid HNO_3
13. hydrochloric acid HCl
14. acetic acid $HC_2H_3O_2$
15. hydrofluoric acid HF
16. phosphorous acid H_3PO_3
17. carbonic acid H_2CO_3
18. nitrous acid HNO_2
19. phosphoric acid H_3PO_4
20. hydrosulfuric acid H_2S

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Answer Key

Name________________

Writing Formulas from Names

Write the formula of each compound.

1. ammonium phosphate — $(NH_4)PO_4$
2. iron(II) oxide — FeO
3. iron(III) oxide — Fe_2O_3
4. carbon monoxide — CO
5. calcium chloride — $CaCl_2$
6. potassium nitrate — KNO_3
7. magnesium hydroxide — $Mg(OH)_2$
8. aluminum sulfate — $Al_2(SO_4)_3$
9. copper(II) sulfate — $CuSO_4$
10. lead(IV) chromate — $Pb(CrO_4)_2$
11. diphosphorus pentoxide — P_2O_5
12. potassium permanganate — $KMnO_4$
13. sodium hydrogen carbonate — $NaHCO_3$
14. zinc nitrate — $Zn(NO_3)_2$
15. aluminum sulfite — $Al_2(SO_3)_3$

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Name________________

Gram Formula Mass

Determine the gram formula mass (the mass of one mole) of each compound.

1. $KMnO_4$ — 158 g
2. KCl — 74.55 g
3. Na_2SO_4 — 142 g
4. $Ca(NO_3)_2$ — 164 g
5. $Al_2(SO_4)_3$ — 342 g
6. $(NH_4)_3PO_4$ — 149 g
7. $CuSO_4 \cdot 5H_2O$ — 250 g
8. $Mg_3(PO_4)_2$ — 262.86 g
9. $Zn(C_2H_3O_2)_2 \cdot 2H_2O$ — 219 g
10. $Zn_3(PO_4)_2 \cdot 4H_2O$ — 458 g
11. H_2CO_3 — 62 g
12. $Hg_2Cr_2O_7$ — 617 g
13. $Ba(ClO_3)_2$ — 304 g
14. $Fe_2(SO_3)_3$ — 352 g
15. $NH_4C_2H_3O_2$ — 77 g

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Name________________

Moles and Mass

Determine the number of moles in each quantity.

1. 25 g of NaCl	0.43 mole
2. 125 g of H_2SO_4	1.28 moles
3. 100. g of $KMnO_4$	0.633 mole
4. 74 g of KCl	0.97 mole
5. 35 g of $CuSO_4 \cdot 5H_2O$	0.14 mole

Determine the number of grams in each quantity.

6. 2.5 mol of NaCl	145 g
7. 0.50 mol of H_2SO_4	49 g
8. 1.70 mol of $KMnO_4$	269 g
9. 0.25 mol of KCl	19 g
10. 3.2 mol of $CuSO_4 \cdot 5H_2O$	800 g

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Name________________

The Mole and Volume

For gases at STP (273 K and 1 atm pressure), one mole occupies a volume of 22.4 L. Identify the volume each quantity of gas will occupy at STP.

1. 1.00 mole of H_2	22.4 L
2. 3.20 moles of O_2	71.7 L
3. 0.750 mole of N_2	16.8 L
4. 1.75 moles of CO_2	39.2 L
5. 0.50 mole of NH_3	11.2 L
6. 5.0 g of H2	56 L
7. 100. g of O_2	70.0 L
8. 28.0 g of N_2	22.4 L
9. 60. g of CO_2	31 L
10. 10. g of NH_3	13 L

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Answer Key

Name____________________

The Mole and Avogadro's Number

One mole of a substance contains Avogadro's number (6.02×10^{23}) of molecules.

Identify how many molecules are in each quantity.

1. 2.0 mol	**1.2×10^{24}**
2. 1.5 mol	**9.0×10^{23}**
3. 0.75 mol	**4.5×10^{23}**
4. 15 mol	**9.0×10^{24}**
5. 0.35 mol	**2.1×10^{23}**

Identify how many moles are in the number of molecules listed.

6. 6.02×10^{23}	**1.00 moles**
7. 1.204×10^{24}	**2.00 moles**
8. 1.5×10^{20}	**2.5×10^{-5} or 0.000025 mole**
9. 3.4×10^{26}	**560 moles**
10. 7.5×10^{19}	**1.2×10^{-4} or 0.00012 mole**

Name____________________

Mixed Mole Problems

Solve each problem.

1. How many grams are there in 1.5×10^{25} molecules of CO_2?
   1.1×10^3 g
2. What volume would the CO_2 in problem 1 occupy at STP?
   5.6×10^2 L
3. A sample of NH_3 gas occupies 75.0 liters at STP. How many molecules is this?
   2.02×10^{24} molecules
4. What is the mass of the sample of NH_3 in problem 3?
   56.9 g
5. How many atoms are there in 1.3×10^{22} molecules of NO_2?
   3.9×10^{22} atoms
6. A 5.0 g sample of O_2 is in a container at STP. What volume is the container?
   3.5 L
7. How many molecules of O_2 are in the container in problem 6? How many atoms of oxygen?
   9.4×10^{22} molecules, 1.9×10^{23} atoms

Name____________________

Percentage Composition

Determine the percentage composition of each compound.

1. $KMnO_4$
   K = **24.7%**
   Mn = **34.8%**
   O = **40.5%**
2. HCl
   H = **2.8%**
   Cl = **97.2%**
3. $Mg(NO_3)_2$
   Mg = **16.2%**
   N = **18.9%**
   O = **64.9%**
4. $(NH_4)PO_4$
   N = **12.4%**
   H = **3.6%**
   P = **27.4%**
   O = **56.6%**
5. $Al_2(SO_4)_3$
   Al = **15.8%**
   S = **28.1%**
   O = **56.1%**

Solve each problem.

6. How many grams of oxygen can be produced from the decomposition of 100. g of $KClO_3$? **39.3 g O**
7. How much iron can be recovered from 25.0 g of Fe_2O_3? **17.5 g Fe**
8. How much silver can be produced from 125 g of Ag_2S? **109 g Ag**

Name____________________

Determining Empirical Formulas

Identify the empirical formula (lowest whole number ratio) of each compound.

1. 75% carbon, 25% hydrogen
   CH_5
2. 52.7% potassium, 47.3% chlorine
   KCl
3. 22.1% aluminum, 25.4% phosphorus, 52.5% oxygen
   $AlPO_4$
4. 13% magnesium, 87% bromine
   $MgBr_2$
5. 32.4% sodium, 22.5% sulfur, 45.1% oxygen
   Na_2SO_4
6. 25.3% copper, 12.9% sulfur, 25.7% oxygen, 36.1% water
   $CuSO_4(H_2O)_5$

Answer Key

Name____________________________

Determining Molecular Formulas (True Formulas)

Solve each problem.

1. The empirical formula of a compound is NO_2. Its molecular mass is 92 g/mol. What is its molecular formula?

   N_2O_4

2. The empirical formula of a compound is CH_2. Its molecular mass is 70 g/mol. What is its molecular formula?

   C_5H_{10}

3. A compound is found to be 40.0% carbon, 6.7% hydrogen and 53.5% oxygen. Its molecular mass is 60. g/mol. What is its molecular formula?

   $C_2H_4O_2$ (CH_3COOH)

4. A compound is 64.9% carbon, 13.5% hydrogen, and 21.6% oxygen. Its molecular mass is 74 g/mol. What is its molecular formula?

   $C_4H_{10}O$ (C_4H_9OH)

5. A compound is 54.5% carbon, 9.1% hydrogen, and 36.4% oxygen. Its molecular mass is 88 g/mol. What is its molecular formula?

   $C_4H_8O_2$ (C_3H_7COOH)

Name____________________________

Composition of Hydrates

A **hydrate** is an ionic compound with water molecules loosely bonded to its crystal structure. The water is in a specific ratio to each formula unit of the salt. For example, the formula $CuSO_4$•$5H_2O$ indicates that there are five water molecules for every one formula unit of $CuSO_4$.

Solve each problem.

1. What percentage of water is found in $CuSO_4$•$5H_2O$?

   35%

2. What percentage of water is found in Na_2S•$9H_2O$?

   67.5%

3. A 5.0 g sample of a hydrate of $BaCl_2$ was heated, and only 4.3 g of the anhydrous salt remained. What percentage of water was in the hydrate?

   14%

4. A 2.5 g sample of a hydrate of $Ca(NO_3)_2$ was heated, and only 1.7 g of the anhydrous salt remained. What percentage of water was in the hydrate?

   32%

5. A 3.0 g sample of Na_2CO_3•H_2O is heated to constant mass. How much anhydrous salt remains?

   2.6 g

6. A 5.0 g sample of $Cu(NO_3)_2$•nH_2O is heated to constant mass. How much anhydrous salt remains?

   3.9 g

Name____________________________

Balancing Chemical Equations

Rewrite and balance each equation.

1. $N_2 + H_2 \rightarrow NH_3$ **$N_2 + 3H_2 \rightarrow 2NH_3$**
2. $KClO_3 \rightarrow KCl + O_2$ **$2KClO_3 \rightarrow 2KCl + 3O_2$**
3. $NaCl + F_2 \rightarrow NaF + Cl_2$ **$2NaCl + F \rightarrow 2NaF + Cl_2$**
4. $H_2 + O_2 \rightarrow H_2O$ **$2H_2 + O_2 \rightarrow 2\ H_2O$**
5. $AgNO_3 + MgCl_2 \rightarrow AgCl + Mg(NO_3)_2$ **$2AgNO_3 + MgCl_2 \rightarrow 2AgCl + Mg(NO_3)_2$**
6. $AlBr_3 + K_2SO \rightarrow KBr + Al_2(SO_4)_3$ **$2AlBr_3 + 3K_2SO_4 \rightarrow 6KBr + Al_2(SO_4)_3$**
7. $CH_4 + O_2 \rightarrow CO_2 + H_2O$ **$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$**
8. $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$ **$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$**
9. $C_8H_{18} + O_2 \rightarrow CO_2 + H_2O$ **$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$**
10. $FeCl_3 + NaOH \rightarrow Fe(OH)_3 + NaCl$ **$FeCl_3 + 3NaOH \rightarrow Fe(OH)_3 + 3NaCl$**
11. $P + O_2 \rightarrow P_2O_5$ **$4P + 5O_2 \rightarrow 2P_2O_5$**
12. $Na + H_2O \rightarrow NaOH + H_2$ **$2Na + 2H_2O \rightarrow 2NaOH + H_2$**
13. $Ag_2O \rightarrow Ag + O_2$ **$2Ag_2O \rightarrow 4Ag + O_2$**
14. $S_8 + O_2 \rightarrow SO_3$ **$S_8 + 12O_2 \rightarrow 8SO_3$**
15. $CO_2 + H_2O \rightarrow C_6H_{12}O_6 + O_2$ **$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$**
16. $K + MgBr_2 \rightarrow KBr + Mg$ **$2K + MgBr_2 \rightarrow 2KBr + Mg$**
17. $HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$ **$2HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$**

Name____________________________

Word Equations

Rewrite each word equation as a chemical equation. Then, balance the equation.

1. zinc + lead(II) nitrate yield zinc nitrate + lead

   $Zn + Pb(NO_3)_2 \rightarrow Zn(NO_3)_2 + Pb$

2. aluminum bromide + chlorine yield aluminum chloride + bromine

   $2AlBr_3 + 3Cl_2 \rightarrow 2AlCl_3 + 3Br_2$

3. sodium phosphate + calcium chloride yield calcium + sodium chloride

   $2Na_3PO_4 + 3CaCl_2 \rightarrow Ca3(PO_4)_2 + 6NaCl$

4. potassium chlorate, when heated, yields potassium chloride + oxygen gas

   $2KClO_3 \rightarrow 2KCl + 3O_2(g)$

5. aluminum + hydrochloric acid yield aluminum chloride + hydrogen gas

   $2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2(g)$

6. calcium hydroxide + phosphoric acid yield calcium phosphate + water

   $3Ca(OH)_2 + 2H_3PO_4 \rightarrow Ca_3(PO_4)_2 + 6H_2O$

7. copper + sulfuric acid yield copper(II) sulfate + water + sulfur dioxide

   $Cu + 2H_2SO_4 \rightarrow CuSO_4 + 2H_2O + SO_2$

8. hydrogen + nitrogen monoxide yield water + nitrogen

   $2H_2 + 2NO \rightarrow 2H_2O + N_2$

Answer Key

Name________________________

Classification of Chemical Reactions

Identify each reaction as *synthesis, decomposition, cationic* or *anionic single replacement*, or *double replacement*.

1. $2H_2 + O_2 \rightarrow 2H_2O$ — **synthesis**
2. $2H_2O \rightarrow 2H_2 + O_2$ — **decomposition**
3. $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$ — **cationic single replacement**
4. $2CO + O_2 \rightarrow 2CO_2$ — **synthesis**
5. $2HgO \rightarrow 2Hg + O_2$ — **decomposition**
6. $2KBr + Cl_2 \rightarrow 2KCl + Br_2$ — **anionic single replacement**
7. $CaO + H_2O \rightarrow Ca(OH)_2$ — **synthesis**
8. $AgNO + NaCl \rightarrow AgCl + NaNO_3$ — **double replacement**
9. $2H_2O_2 \rightarrow 2H_2O + O_2$ — **decomposition**
10. $Ca(OH)_2 + H_2SO_4 \rightarrow CaSO_4 + 2H_2O$ — **double replacement**

Name________________________

Predicting Products of Chemical Reactions

Predict the product in each reaction. Then, write the balanced equation and classify the reaction.

Predictions will vary.

1.	magnesium bromide + chlorine **$MgBr_2 + Cl_2 \longrightarrow MgCl_2Br_2$**	**anionic single replacement**
2.	aluminum + iron(III) oxide **$2Al + Fe_2O_3 \longrightarrow 2Fe + Al_2O_3$**	**cationic single replacement**
3.	silver nitrate + zinc chloride **$2AgNO_3 + ZnCl_2 \longrightarrow 2AgCl + Zn(NO_3)_2$**	**double replacement**
4.	hydrogen peroxide (catalyzed by manganese dioxide) **$2H_2O_2 \xrightarrow{MnO_2} 2H_2O + O_2$**	**decomposition**
5.	zinc + hydrochloric acid **$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$**	**cationic single replacement**
6.	sulfuric acid + sodium hydroxide **$H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$**	**double replacement (neutralization)**
7.	sodium + hydrogen **$2Na + H_2 \longrightarrow 2NaOH$**	**synthesis**
8.	acetic acid + copper **CH_3COOH (or $HC_2H_3O_2$) $+ Cu \longrightarrow$ no reaction**	**none**

Name________________________

Stoichiometry: Mole-Mole Problems

Solve each problem.

1. $N_2 + 3H_2 \rightarrow 2NH_3$
   How many moles of hydrogen are needed to completely react with two moles of nitrogen?
   6 moles
2. $2KClO_3 \rightarrow 2KCl + 3O_2$
   How many moles of oxygen are produced by the decomposition of six moles of potassium chlorate?
   9 moles
3. $Zn + 2HCl \rightarrow ZnCl_2 + H_2$
   How many moles of hydrogen are produced from the reaction of three moles of zinc with an excess of hydrochloric acid?
   3 moles
4. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
   How many moles of oxygen are necessary to react completely with four moles of propane (C_3H_8)?
   20 moles
5. $K_3PO_4 + Al(NO_3)_3 \rightarrow 3KNO_3 + AlPO_4$
   How many moles of potassium nitrate (KNO_3) are produced when six moles of potassium phosphate (KPO_4) react with two moles of aluminum nitrate($Al(NO_3)_3$)?
   6 moles

Name________________________

Stoichiometry: Volume-Volume Problems

Solve each problem.

1. $N_2 + 3H_2 \rightarrow 2NH_3$
   What volume of hydrogen is necessary to react with five liters of nitrogen to produce ammonia? (Assume constant temperature and pressure.)
   15 L
2. What volume of ammonia is produced in the reaction in problem 1?
   10 L
3. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
   If 20 liters of oxygen are consumed in the above reaction, how many liters of carbon dioxide are produced?
   12 L
4. $2H_2O \rightarrow 2H + O_2$
   If 30 mL of hydrogen are produced in the above reaction, how many milliliters of oxygen are produced?
   15 mL
5. $2CO + O_2 \rightarrow 2CO_2$
   How many liters of carbon dioxide are produced if 75 liters of carbon monoxide are burned in oxygen? How many liters of oxygen are necessary?
   75 L CO_2
   37.5 L O_2

Answer Key

Name________________________

Stoichiometry: Mass-Mass Problems

Solve each problem.

1. $2KClO_3 \rightarrow 2KCl + 3O_2$
   How many grams of potassium chloride are produced if 25 g of potassium chlorate decompose?
   15 g
2. $N_2 + 3H_2 \rightarrow 2NH_3$
   How many grams of hydrogen are necessary to react completely with 50.0 g of nitrogen in the above reaction?
   10.7 g
3. How many grams of ammonia are produced in the reaction in problem 2?
   60.7 g
4. $2AgNO_3 + BaCl \rightarrow 2AgCl + Ba(NO_3)$
   How many grams of silver chloride are produced from 5.0 g of silver nitrate reacting with an excess of barium chloride?
   4.2 g
5. How much barium chloride is necessary to react with the silver nitrate in problem 4?
   3.1 g

Name________________________

Stoichiometry: Mixed Problems

Solve each problem.

1. $N_2 + 3H_2 \rightarrow 2NH_3$
   What volume of NH_3 at STP is produced if 25.0 g of N_2 is reacted with an excess of H_2?
   40.0 L
2. $2KClO_3 \rightarrow 2KCl + 3O_2$
   If 5.0 g of $KClO_3$ are decomposed, what volume of O_2 is produced at STP?
   1.4 L
3. How many grams of KCl are produced in problem 2?
   3.0 g
4. $Zn + 2HCl \rightarrow ZnCl_2 + H_2$
   What volume of hydrogen at STP is produced when 2.5 g of zinc react with an excess of hydrochloric acid?
   0.86 L
5. $H_2SO_4 + 2NaOH \rightarrow H_2O + Na_2SO_4$
   How many molecules of water are produced if 2.0 g of sodium sulfate are produced in the above reaction?
   8.5×10^{21} molecules
6. $2AlCl_3 \rightarrow 2Al + 3Cl_2$
   If 10.0 g of aluminum chloride are decomposed, how many molecules of Cl_2 are produced?
   6.77×10^{22} molecules

Name________________________

Stoichiometry: Limiting Reagent

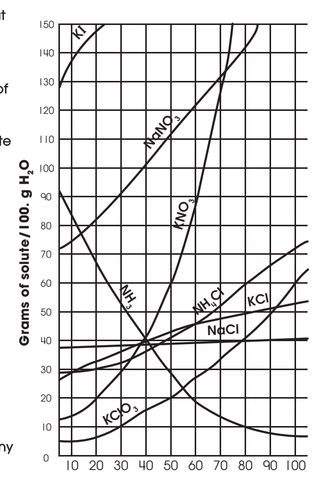
Solve each problem.

1. $N_2 + 3H_2 \rightarrow 2NH_3$
   How many grams of NH_3 can be produced from the reaction of 28 g of N_2 and 25 g of H_2?
   34 g
2. How much of the excess reagent in problem 1 is left over?
   19 g
3. $Mg + 2HCl \rightarrow MgCl_2 + H_2$
   What volume of hydrogen at STP is produced from the reaction of 50.0 g of Mg and the equivalent of 75 g of HCl?
   23.0 L
4. How much of the excess reagent in problem 3 is left over?
   24.3 g
5. $3AgNO_3 + Na_3PO_4 \rightarrow Ag_3PO_4 + 3NaNO_3$
   Silver nitrate and sodium phosphate are reacted in equal amounts of 200. g each. How many grams of silver phosphate are produced?
   165 g
6. How much of the excess reagent in problem 5 is left over?
   136 g

Name________________________

Solubility Curves

Answer each question based on the solubility curve shown.



1. Which salt is least soluble in water at 20°C? **$KClO_3$**
2. How many grams of potassium chloride can be dissolved in 200 g of water at 80°C? **100 g**
3. At 40°C, how much potassium nitrate can be dissolved in 300 g of water? **126 g**
4. Which salt shows the least change in solubility from 0°C to 100°C? **NaCl**
5. At 30°C, 85 g of sodium nitrate are dissolved in 100 g of water. Is this solution *saturated*, *unsaturated*, or *supersaturated*? **unsaturated**
6. A saturated solution of potassium chlorate is formed from 100 g of water. If the saturated solution is cooled from 80°C to 50°C, how many grams of precipitate are formed? **7 g**
7. What compound shows a decrease in solubility from 0°C to 100°C? **NH_3**
8. Which salt is most soluble at 10°C? **KI**
9. Which salt is least soluble at 50°C? **$KClO_3$**
10. Which salt is least soluble at 90°C? **NH_3**

Answer Key

Name____________________

Molarity (M)

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liter of solution}}$$

Solve each problem.

1. What is the molarity of a solution in which 58 g of NaCl are dissolved in 1.0 L of solution?

   1.0 M

2. What is the molarity of a solution in which 10.0 g of $AgNO_3$ are dissolved in 500. mL of solution?

   0.118 M

3. How many grams of KNO_3 should be used to prepare 2.00 L of a 0.500 M solution?

   101 g

4. To what volume should 5.0 g of KCl be diluted in order to prepare a 0.25 M solution?

   270 mL

5. How many grams of $CuSO_4 \bullet 5H_2O$ are needed to prepare 100. mL of a 0.10 M solution?

   2.5 g

Name____________________

Molarity by Dilution

Acids are usually acquired from chemical supply houses in concentrated form. These acids are diluted to the desired concentration by adding water. Since moles of acid before dilution equal moles of acid after dilution, and moles of acid = $M \times V$, then $M_1 \times V_1 = M_2 \times V_2$.

Solve each problem.

1. How much concentrated 18 M sulfuric acid is needed to prepare 250 mL of a 6.0 M solution?

   83 mL

2. How much concentrated 12 M hydrochloric acid is needed to prepare 1$\bar{0}$0 mL of a 2.0 M solution?

   17 mL

3. To what volume should 25 mL of 15 M nitric acid be diluted to prepare a 3.0 M solution?

   125 mL

4. How much water should be added to 50. mL of 12 M hydrochloric acid to produce a 4.0 M solution?

   100 mL (150 mL total solution)

5. How much water should be added to100. mL of 18 M sulfuric acid to prepare a 1.5 M solution?

   1.1 L (1.2 L or 1,200 mL total solution)

Name____________________

Molality (m)

$$\text{Molality} = \frac{\text{moles of solute}}{\text{kg of solvent}}$$

Solve each problem.

1. What is the molality of a solution in which 3.0 moles of NaCl are dissolved in 1.5 kg of water?

   2.0 m

2. What is the molality of a solution in which 25 g of NaCl are dissolved in 2.0 kg of water?

   0.22 m

3. What is the molality of a solution in which 15 g of I_2 are dissolved in 500. g of alcohol?

   0.12 m

4. How many grams of I_2 should be added to 750 g of CCl_4 to prepare a 0.020 m solution?

   3.8 g

5. How much water should be added to 5.00 g of KCl to prepare a 0.500 m solution?

   135 g

Name____________________

Normality (N)

normality = molarity × total positive oxidation number of solute

Example: What is the normality of 3.0 M of H_2SO_4?

Since the total positive oxidation number of H_2SO_4 is +2 (2 H^+), N = 6.0.

Solve each problem.

1. What is the normality of a 2.0 M NaOH solution?

   2.0 N

2. What is the normality of a 2.0 M H_3PO_4 solution?

   6.0 N

3. A solution of H_2SO_4 is 3.0 N. What is its molarity?

   1.5 M

4. What is the normality of a solution in which 2.0 g of $Ca(OH)_2$ is dissolved in 1.0 L of solution?

   0.054 N

5. How much $AlCl_3$ should be dissolved in 2.00 L of solution to produce a 0.150 N solution?

   13.3 g

Answer Key

Name________________________

Electrolytes

Electrolytes are substances that break up (dissociate or ionize) in water to produce ions. These ions are capable of conducting an electric current.

Generally, electrolytes consist of acids, bases, and salts (ionic compounds). Nonelectrolytes are usually covalent compounds, with the exception of acids.

Check the appropriate column to classify each compound as either an electrolyte or a nonelectrolyte.

Compound	Electrolyte	Nonelectrolyte
1. NaCl	✓	
2. CH_3OH (methyl alcohol)		✓
3. $C_3H_5(OH)_3$ (glycerol)		✓
4. HCl	✓	
5. $C_6H_{12}O_6$ (sugar)		✓
6. NaOH	✓	
7. C_2H_5OH (ethyl alcohol)		✓
8. CH_3COOH (acetic acid)	✓	
9. NH_4OH ($NH_3 + H_2O$)	✓	
10. H_2SO_4	✓	

Name________________________

Effect of a Solute on Freezing and Boiling Points

Use the following formulas to calculate changes in freezing and boiling points due to the presence of a nonvolatile solute. The freezing point is always lowered; the boiling point is always raised.

$\Delta T_F = M \times d.f. \times k_F$ $\qquad k_BH_2O = 0.52°C/M$

$\Delta T_B = M \times d.f. \times k_B$ $\qquad k_FH_2O = 1.86°C/M$

M = molality of solution

k_F and k_B = constants for particular solvent

$d.f.$ = dissociation factor (how many particles the solute breaks up into; for a nonelectrolyte, $d.f.$ = 1)

(The theoretical dissociation factor is always greater than observed effect.)

Solve each problem.

1. What is the new boiling point if 25 g of NaCl are dissolved in 1.0 kg of water?

100.45°C

2. What is the freezing point of the solution in problem 1?

–1.6°C

3. What are the new freezing and boiling points of water if 50. g of ethylene glycol (molecular mass = 62 g/M) are added to 50. g of water?

boiling point: 108.4°C **freezing point: –30°C**

4. When 5.0 g of a nonelectrolyte are added to 25 g of water, the new freezing point is -2.5°C. What is the molecular mass of the unknown compound?

149 g

Name________________________

Solubility (Polar vs. Nonpolar)

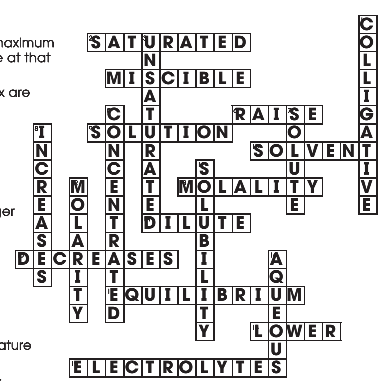
Generally, "like dissolves like." Polar molecules dissolve other polar molecules and ionic compounds. Nonpolar molecules dissolve other nonpolar molecules. Alcohols, which have characteristics of both, tend to dissolve in both types of solvents but will not dissolve ionic solids.

Check the appropriate columns to indicate whether the solute is soluble in a polar or nonpolar solvent.

Solutes	Solvents: water	Solvents: CCl_4	Solvents: alcohol
1. NaCl	✓		
2. I_2		✓	✓
3. ethanol	✓	✓	✓
4. benzene		✓	✓
5. Br_2		✓	✓
6. KNO_3	✓		
7. toluene		✓	✓
8. $Ca(OH)_2$	✓		

Name________________________

Solutions Crossword



Across

2. Solution containing the maximum amount of solute possible at that temperature
4. Two liquids which can mix are said to be ___________.
6. The presence of a nonvolatile solute will ___________ the boiling point of a solvent.
9. A homogeneous mixture
10. Substance present in larger amounts in a mixture
13. Moles of a solute per kilogram of solvent
14. Solution containing a relatively large amount of solvent
15. The solubility of gases ___________ as temperature increases.
17. State in which the rate of dissolving is equal to the rate of precipitation
18. The presence of a nonvolatile solute will ___________ the freezing point of a solvent.
19. These substances dissociate or ionize in water and are then able to conduct an electric current.

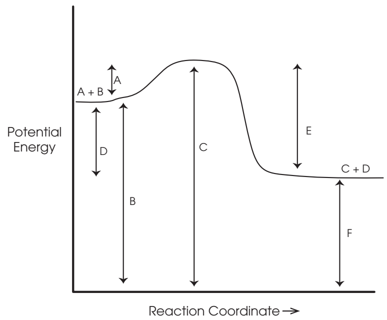
Down

1. Properties that depend on the number of particles in a solution
3. Solution in which more solute can be dissolved
5. Solution containing a relatively large amount of dissolved solute
7. Substance present in a smaller amount in a mixture
8. The solubility of most solids ___________ as temperature increases.
11. Maximum amount of solute that can dissolve in a stated amount of solute at a given temperature
12. Moles of solute per liter of solution
16. Solutions in which water is the solvent are called ___________.

Answer Key

Name____________________

Potential Energy Diagram



A + B ⟷ C + D + energy

Answer each question using the graph shown.

1. Is the above reaction endothermic or exothermic? **exothermic**
2. Which letter represents the potential energy of the reactants? **B**
3. Which letter represents the potential energy of the products? **F**
4. Which letter represents the heat of reaction (ΔH)? **D**
5. Which letter represents the activation energy of the forward reaction? **A**
6. Which letter represents the activation energy of the reverse reaction? **E**
7. Which letter represents the potential energy of the activated complex? **C**
8. Is the reverse reaction endothermic or exothermic? **endothermic**
9. If a catalyst were added, what letter(s) would change? **A, C, E**

Name____________________

Entropy

Entropy is the degree of randomness in a substance. The symbol for change in entropy is ΔS.
Solids are very ordered and have low entropy. Liquids and aqueous ions have more entropy because they move about more freely. Gases have an even larger amount of entropy. According to the Second Law of Thermodynamics, nature is always proceeding to a state of higher entropy.

Determine whether each reaction shows an *increase* or *decrease* in entropy.

1. $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$ **increase**
2. $H_2O(l) \rightarrow H_2O(s)$ **decrease**
3. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ **decrease**
4. $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ **increase**
5. $KCl(s) \rightarrow KCl(l)$ **increase**
6. $CO_2(s) \rightarrow CO_2(g)$ **increase**
7. $H^+(aq) + C_2H_3O_2^-(aq) \rightarrow HC_2H_3O_2(l)$ **decrease**
8. $C(s) + O_2(g) \rightarrow CO_2(g)$ **increase**
9. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ **no change**
10. $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$ **increase**
11. $2N_2O_5(g) \rightarrow 4NO_2(g) + O_2(g)$ **increase**
12. $2Al(s) + 3I_2(s) \rightarrow 2AlI_3(s)$ **decrease**
13. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ **decrease**
14. $2NO(g) \rightarrow N_2(g) + O_2(g)$ **no change**
15. $H_2O(g) \rightarrow H_2O(l)$ **decrease**

Name____________________

Gibbs Free Energy

For a reaction to be spontaneous, the sign of ΔG (Gibbs free energy) must be negative. The mathematical formula for this value is:

$$\Delta G = \Delta H - T\Delta S$$

where ΔH = change in enthalpy or heat of reaction
T = temperature in Kelvin
ΔS = change in entropy or randomness

Complete the table for the sign of ΔG: +, –, or *undetermined*. When conditions allow for an undetermined sign of ΔG, temperature will decide spontaneity.

ΔH	ΔS	ΔG
–	+	**–**
+	–	**+**
–	–	**undetermined**
+	+	**undetermined**

Answer each question.

1. The conditions in which ΔG is always negative are when ΔH is **negative** and ΔS is **positive**.
2. The conditions in which ΔG is always positive are when ΔH is **positive** and ΔS is **negative**.
3. When the situation is indeterminate, a low temperature favors the (entropy, enthalpy) factor and a high temperature favors the (entropy, enthalpy) factor.

Answer problems 4–6 with *always*, *sometimes*, or *never*.

4. The reaction: $Na(OH)_s \rightarrow Na^+(aq) + OH^-(aq)$ + energy will **always** be spontaneous.
5. The reaction: energy + $2H_2(g) + O_2(g) + 2H_2O(l)$ will **never** be spontaneous.
6. The reaction: energy + $H_2O(s) \rightarrow H_2O(l)$ will **sometimes** be spontaneous.
7. What is the value of ΔG if ΔH = 32.0 kJ, ΔS = +25.0 kJ/K and T = 293 K? **–7,293 kJ**
8. Is the reaction in problem 7 spontaneous? **yes**
9. What is the value of ΔG if ΔH = +12.0 kJ, ΔS = 5.00 kJ/K and T = 290. K? **+1,460 kJ**
10. Is the reaction in problem 9 spontaneous? **no**

Name____________________

Equilibrium Constant (K)

Write the expression for the equilibrium constant (K) for each reaction.

1. $N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$

$$K = \frac{(NH_3)^2}{(N_2)(H_2)^3}$$

2. $2KClO_3(s) \leftrightarrow 2KCl(s) + 3O_2(g)$

$$K = (O_2)^3$$

3. $H_2O(l) \leftrightarrow H^+(aq) + OH^-(aq)$

$$K = (H^+)(OH^-)$$

4. $2CO(g) + O_2(g) \leftrightarrow 2CO_2(g)$

$$K = \frac{(CO_2)^2}{(CO)^2(O_2)}$$

5. $Li_2CO_3(s) \rightarrow 2Li^+(aq) + CO_3^{2-}(aq)$

$$K = (Li^+)^2(CO_3^{-2})$$

Answer Key

Name_________________________

Calculations Using the Equilibrium Constant

Using the equilibrium constant expressions you determined on page 80, calculate the value of K when:

1. $(NH_3) = 0.0100$ M, $(N_2) = 0.0200$ M, $(H_2) = 0.0200$ M

$$\frac{(0.01)^2}{(0.02)(0.02)^3} = 625$$

2. $(O_2) = 0.0500$ M

$$(0.05)^3 = 1.25 \times 10^{-4}$$

3. $(H^+) = 1 \times 10^{-8}$ M, $(OH^-) = 1 \times 10^{-6}$ M

$$(1 \times 10^{-8})(1 \times 10^{-6}) = 1 \times 10^{-14}$$

4. $(CO) = 2.0$ M, $(O_2) = 1.5$ M, $(CO_2) = 3.0$ M

$$\frac{(3.0)^2}{(2.0)^2(1.5)} = 1.5$$

5. $(Li^+) = 0.2$ M, $(CO_3^{-2}) = 0.1$ M

$$(0.2)^2(0.1) = 4 \times 10^{-3}$$

Name_________________________

Le Chatelier's Principle

Le Chatelier's principle states that when a system at equilibrium is subjected to a stress, the system will shift its equilibrium point in order to relieve the stress.

Complete the chart by writing ***left***, ***right***, or ***none*** for equilibrium shift. Then, write ***decreases***, ***increases***, or ***remains the same*** for the concentrations of reactants and products, and for the value of K. The first one has been done for you.

$$N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g) + 22.0 \text{ kcal}$$

Stress	Equilibrium Shift	(N_2)	(H_2)	(NH_3)	K
1. Add N_2	right	———	decreases	increases	remains the same
2. Add H_2	**right**	**decreases**	———	**increases**	**same**
3. Add NH_3	**left**	**increases**	**increases**	———	**same**
4. Remove N_2	**left**	———	**increases**	**decreases**	**same**
5. Remove H_2	**left**	**increases**	———	**decreases**	**same**
6. Remove NH_3	**right**	**decreases**	**decreases**	———	**same**
7. Increase temperature	**left**	**increases**	**increases**	**decreases**	**increases**
8. Decrease temperature	**right**	**decreases**	**decreases**	**increases**	**decreases**
9. Increase pressure	**right**	**decreases**	**decreases**	**increases**	**same**
10. Decrease pressure	**left**	**increases**	**increases**	**decreases**	**same**

Name_________________________

Le Chatelier's Principle (Cont.)

$$12.6 \text{ kcal} + H_2(g) + I_2(g) \leftrightarrow 2HI(g)$$

Stress	Equilibrium Shift	(H_2)	(I_2)	(HI)	K
11. Add H_2	right	———	decreases	increases	remains the same
12. Add I_2	**right**	**decreases**	———	**increases**	**same**
13. Add HI	**left**	**increases**	**increases**	———	**same**
14. Remove H_2	**left**	———	**increases**	**decreases**	**same**
15. Remove I_2	**left**	**increases**	———	**decreases**	**same**
16. Remove HI	**right**	**decreases**	**decreases**	———	**same**
17. Increase temperature	**right**	**decreases**	**decreases**	**increases**	**increases**
18. Decrease temperature	**left**	**increases**	**increases**	**decreases**	**decreases**
19. Increase pressure	**none**	**same**	**same**	**same**	**same**
20. Decrease pressure	**none**	**same**	**same**	**same**	**same**

$$NAOH(s) \leftrightarrow Na^+(aq) + OH^-(aq) + 10.6 \text{ kcal}$$

Stress	Equilibrium Shift	Amount NaOH(s)	(Na^+)	(OH^-)	K
21. Add NaOH(s)	**none**	———	**same**	**same**	**same**
22. Add NaCl (Adds Na^+)	**left**	**increases**	———	**decreases**	**same**
23. Add KOH (Adds OH^-)	**left**	**increases**	**decreases**	———	**same**
24. Add H^+ (Removes OH^-)	**right**	**decreases**	**increases**	———	**same**
25. Increase temperature	**left**	**increases**	**decreases**	**decreases**	**decreases**
26. Decrease temperature	**right**	**decreases**	**increases**	**increases**	**increases**
27. Increase pressure	**none**	**same**	**same**	**same**	**same**
28. Decrease pressure	**none**	**same**	**same**	**same**	**same**

Name_________________________

Bronsted–Lowry Acids and Bases

According to **Bronsted–Lowry theory**, an acid is a proton (H^+) donor and a base is a proton acceptor.

Example: $HCl + OH^- \xrightarrow{H^+} Cl^- + H_2O$

The HCl acts as an acid, and the OH^- acts as a base. This reaction is reversible in that the H_2O can give back the proton to the Cl^-.

Label the Bronsted-Lowry acids and bases in each reaction and show the direction of proton transfer.

Example: H_2O (acid) + Cl^- (base) $\leftrightarrow$ OH^- (base) + HCl (acid)
H+ H+

1. $H_2O + H_2O \leftrightarrow H_3O^+ + OH^-$

H+ H+
B A A B

2. $H_2SO_4 + OH^- \leftrightarrow HSO_4^- + H_2O$

H+ H+
A B B A

3. $HSO_4^- + H_2O \leftrightarrow SO_4^{2-} + H_3O^+$

A B B A

4. $OH^- + H_3O+ \leftrightarrow H_2O + H_2O$

H+ H+
B A A B

5. $NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$

H+ H+
B A A B

Answer Key

Name

Conjugate Acid–Base Pairs

In the exercise on page 84, it was shown that after an acid has given up its proton, it is capable of getting the proton back and acting as a base. **Conjugate base** is what is left after an acid gives up a proton. The stronger the acid, the weaker the conjugate base. The weaker the acid, the stronger the conjugate base.

Complete the chart.

Conjugate Pairs

	Acid	Base	Equation
1.	H_2SO_4	HSO_4^-	$H_2SO_4 \longleftrightarrow H^+ + HSO_4^-$
2.	H_3PO_4	**$H_2PO_4^-$**	**$H_3PO_4 \rightarrow H^+ + H_2PO_4^-$**
3.	**HF**	F^-	**$HF \rightarrow H^+ + F^-$**
4.	**HNO_3**	NO_3^-	**$HNO_3 \longleftrightarrow H^+ + NO_3^-$**
5.	$H_2PO_4^-$	**HPO_4^{-2}**	**$H_2PO_4^- \longleftrightarrow H^+$**
6.	H_2O	**OH^-**	**$H_2O \longleftrightarrow H^+ + OH^-$**
7.	**HSO_4^-**	SO_4^{2-}	**$HSO_4^- \longleftrightarrow H^+ + SO_4^{-3}$**
8.	HPO_4^{-2}	**PO_4^{-3}**	**$HPO_4^{-2} \longleftrightarrow H^+ + PO_4^{-3}$**
9.	NH_4^+	**NH_3**	**$NH_4^+ \longleftrightarrow H^+ + NH_3$**
10.	**H_3O^+**	H_2O	**$H_3O \longleftrightarrow H^+ + H_2O$**

11. Which is a stronger base, HSO_4^- or $H_2PO_4^-$? **$H_2PO_4^-$**

12. Which is a weaker base, Cl^- or NO_2^- ? **Cl^-**

Name

pH and pOH

The pH of a solution indicates how acidic or basic that solution is. A pH of less than 7 is acidic; a pH of 7 is neutral; and a pH greater than 7 is basic.

Since $[H^+][OH^-] = 10^{-14}$ at 25°C, if $[H^+]$ is known, the $[OH^-]$ can be calculated and vice versa.

$pH = -\log [H^+]$ So if $[H^+] = 10^{-6}$ M, pH = 6.

$pOH = -\log [OH^-]$ So if $[OH^-] = 10^{-8}$ M, pOH = 8.

Together, pH + pOH = 14.

Complete the chart.

	$[H^+]$	pH	$[OH^-]$	pOH	Acidic or Basic
1.	10^{-5} M	5	10^{-9} M	9	acidic
2.	**10^{-7} M**	7	**10^{-7} M**	**7**	**neutral**
3.	**10^{-10} M**	**10**	10^{-4} M	**14**	**basic**
4.	10^{-2} M	**2**	**10^{-12} M**	**12**	**acidic**
5.	**10^{-3} M**	**3**	**10^{-11} M**	11	**acidic**
6.	**10^{-12} M**	12	**10^{-2} M**	**2**	**basic**
7.	**10^{-9} M**	**9**	10^{-5} M	**5**	**basic**
8.	10^{-11} M	**11**	**10^{-3} M**	**3**	**basic**
9.	**10^{-1} M**	**1**	**10^{-13} M**	13	**acidic**
10.	**10^{-6} M**	6	**10^{-8} M**	**8**	**acidic**

Name

pH of Solutions

Calculate the pH of each solution.

1. 0.01 M HCl	**pH = 2**
2. 0.0010 M NaOH	**pH = 11**
3. 0.050 M $Ca(OH)_2$	**pH = 1.3**
4. 0.030 M HBr	**pH = 1.5**
5. 0.150 M KOH	**pH = 13.2**
6. 2.0 M $HC_2H_3O_2$ (Assume 5.0% dissociation.)	**pH = 1.0**
7. 3.0 M HF (Assume 10.0% dissociation.)	**pH = 0.52**
8. 0.50 M HNO_3	**pH = 0.30**
9. 2.50 M NH_4OH (Assume 5.00% dissociation.)	**pH = 13.1**
10. 5.0 M HNO_2 (Assume 1.0% dissociation.)	**pH = 1.3**

Name

Acid–Base Titration

To determine the concentration of an acid (or base), we can react it with a base (or acid) of known concentration until it is completely neutralized. This point of exact neutralization, known as the **endpoint**, is noted by the change in color of the indicator. Use the following equation:

$$N_A \times V_A = N_B \times V_B$$

where N = normality
V = volume

Solve each problem.

1. A 25.0 mL sample of HCl was titrated to the endpoint with 15.0 mL of 2.0 N NaOH. What was the normality of the HCl? **1.2 N**
2. A 10.0 mL sample of H_2SO_4 was exactly neutralized by 13.5 mL of 1.0 M KOH. What is the normality of the H_2SO_4? **1.4 N**
3. How much 1.5 M NaOH is necessary to exactly neutralize 20.0 mL of 2.5 M H_3PO_4? **33 mL**
4. How much 0.5 M HNO_3 is necessary to titrate 25.0 mL of 0.05 M $Ca(OH)_2$ solution to the endpoint? **2.5 mL**
5. What is the molarity of a NaOH solution if 15.0 mL is exactly neutralized by 7.5 mL of a 0.02 M $HC_2H_3O_2$ solution? **0.01 M**

Answer Key

Name________________________

Hydrolysis of Salts

Salt solutions may be acidic, basic, or neutral, depending on the original acid and base that formed the salt.

strong acid + strong base → neutral salt
strong acid + weak base → acidic salt
weak acid + strong base → basic salt

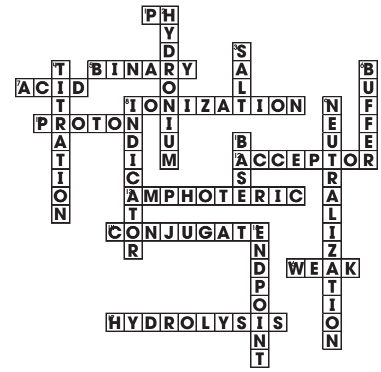
A weak acid and a weak base will produce any type of solution depending on the relative strengths of the acid and base involved.

Complete the chart for each salt shown.

Salt	Parent Acid	Parent Base	Type of Solution
1. KCl	**HCl**	**KOH**	**neutral**
2. NH_4NO_3	**HNO_3**	**NH_4OH ($NH_3 + H_2O$)**	**acidic**
3. Na_3PO_4	**H_3PO_4**	**NaOH**	**basic**
4. $CaSO_4$	**H_2SO_4**	**$Ca(OH)_2$**	**neutral**
5. $AlBr_3$	**HBr**	**$Al(OH)_2$**	**acidic**
6. CuI_2	**HI**	**$Cu(OH)_2$**	**acidic**
7. MgF_4	**$Mg(OH)_2$**	**HF**	**basic**
8. $NaNO_3$	**HNO_3**	**NaOH**	**neutral**
9. $LiC_2H_3O_2$	**$HC_2H_3O_2$**	**LiOH**	**basic**
10. $ZnCl_2$	**HCl**	**$Zn(OH)_2$**	**acidic**
11. $SrSO_4$	**H_2SO_4**	**$Sr(OH)_2$**	**neutral**
12. $Ba_3(PO_4)_2$	**H_3PO_4**	**$Ba(OH)_2$**	**basic**

Name________________________

Acids and Bases Crossword



Across

1. Scale of acidity
5. An acid that consists of only two elements
7. Substance that forms hydronium ions in water (Arrhenius __________)
8. This happens when an acid dissolves in water.
10. According to Bronsted-Lowry, an acid is a ____________ donor.
12. According to Bronsted-Lowry, a base is a proton ____________.
13. Can act as either an acid or a base
14. These pairs differ only by a proton.
16. An acid with a small K_a value would be a ____________ acid.
17. Reaction of an ion with H_2O to produce $H^+(aq) + OH^-(aq)$

Down

2. H_3O^+
3. Formed from the reaction of an acid and a base
4. Procedure to determine the concentration of an acid or base
6. A solution that will resist changes in pH
8. Changes color at the endpoint of a titration
9. The reaction of an acid with a base
11. Substance that produces hydroxide ions in aqueous solution (Arrhenius __________)
15. When equivalent amounts of H^+ and OH^- have reacted in a titration

Name________________________

Assigning Oxidation Numbers

Assign oxidation numbers to all of the elements in each compound or ion shown.

1. HCl $\overset{+1}{H}\overset{-1}{Cl}$	11. H_2SO_3 $\overset{+1}{H_2}\overset{+4}{S}\overset{-2}{O_3}$
2. KNO_3 $\overset{+1}{K}\overset{+5}{N}\overset{-2}{O_3}$	12. H_2SO_4 $\overset{+1}{H_2}\overset{+6}{S}\overset{-2}{O_4}$
3. OH^- $\overset{-2}{O}\overset{+1}{H^-}$	13. BaO_2 $\overset{+2}{Ba}\overset{-1}{O_2}$
4. Mg_3N_2 $\overset{+2}{Mg_3}\overset{-3}{N_2}$	14. $KMnO_4$ $\overset{+1}{K}\overset{+7}{Mn}\overset{-2}{O_4}$
5. $KClO_3$ $\overset{+1}{K}\overset{+5}{Cl}\overset{-2}{O_3}$	15. LiH $\overset{+1}{Li}\overset{-1}{H}$
6. $Al(NO_3)_3$ $\overset{+3}{Al}(\overset{+5}{N}\overset{-2}{O_3})_3$	16. MnO_2 $\overset{+4}{Mn}\overset{-2}{O_2}$
7. S_8 $\overset{0}{S_8}$	17. OF_2 $\overset{+2}{O}\overset{-1}{F_2}$
8. H_2O_2 $\overset{+1}{H_2}\overset{-1}{O_2}$	18. SO_3 $\overset{+6}{S}\overset{-2}{O_3}$
9. PbO_2 $\overset{+4}{Pb}\overset{-2}{O_2}$	19. NH_3 $\overset{-3}{N}\overset{+1}{H_3}$
10. $NaHSO_4$ $\overset{+1}{Na}\overset{+1}{H}\overset{+6}{S}\overset{-2}{O_4}$	20. Na $\overset{0}{Na}$

Name________________________

Redox Reactions

For each equation, identify the substance oxidized, the substance reduced, the oxidizing agent, and the reducing agent. Then, write the oxidation and reduction half-reactions.

Example: oxidized: Mg, reduced: Br_2

$Mg + Br_2 \rightarrow MgBr_2$

reducing agent: Mg; oxidizing agent: Br_2

oxidation half-reaction: $Mg^0 \rightarrow Mg^{+2} + 2e^-$
reduction half-reaction: $2e^- + Br_2^0 \rightarrow 2Br^-$

oxidized **reduced**
1. $2H_2 + O_2 \rightarrow 2H_2O$
reducing agent **oxidizing agent**

oxidation half-reaction: $2H_2^0 \rightarrow 4H^+ + 4e^-$
reduction half-reaction: $4e^- + O_2^0 \rightarrow 2O^{2-}$

oxidized **reduced**
2. $Fe + Zn^{2+} \rightarrow Fe^{2+} + Zn$
reducing agent **oxidizing agent**

oxidation half-reaction: $Fe \rightarrow Fe^{2+} + 2e^-$
reduction half-reaction: $2e^- + Zn^{2+} \rightarrow Zn^0$

oxidized **reduced**
3. $2Al + 3Fe^{2+} \rightarrow 2Al^{3+} + 3Fe$
reducing agent **oxidizing agent**

oxidation half-reaction: $2Al^0 \rightarrow 2Al^{3+} + 6e^-$
reduction half-reaction: $6e^- + 3Fe^{2+} \rightarrow 3Fe^0$

oxidized **reduced**
4. $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$
reducing agent **oxidizing agent**

oxidation half-reaction: $Cu \rightarrow Cu^{2+} + 2e^-$
reduction half-reaction: $2e^- + 2Ag^+ \rightarrow 2Ag$

Answer Key

Name________________

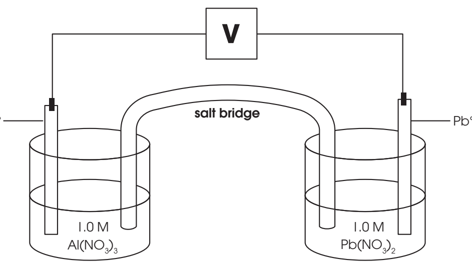
Balancing Redox Equations

Balance each equation using the half-reaction method.

1. Sn° + **2** Ag^{+} → Sn^{2+} + **2** Ag°	**oxidizing: $Sn^{0} \rightarrow Sn^{+2} + 2e^{-}$** **reduced: $2(e^{-} + Ag^{+} \rightarrow Ag0)$**
2. **2** Cr° + **3** Pb^{2+} → **2** Cr^{3+} + **3** Pb°	**oxidizing: $2(Cr^{0} \rightarrow Cr^{+3} + 3e^{-}$** **reducing: $3(2e^{-} + Pb^{2+} \rightarrow Pb^{0})$**
3. **2** $KClO_3$ → **2** KCl + **3** O_2	**oxidizing: $2O_3^{2+} \rightarrow 3O_2^{0} + 12e^{-}$** **reducing: $2(6e^{-} + Cl^{5+} \rightarrow Cl^{-})$**
4. **4** NH_3 + **5** O_2 → **4** NO + **6** H_2O	**oxidizing: $4(N^{3-} \rightarrow N^{2+} + 5e^{-})$** **reducing: $5(e^{-} + O_2 \rightarrow 2O^{2-})$**
5. PbS + **4** H_2O_2 → $PbSO_4$ + **4** H_2O	**oxidizing: $S^{2-} \rightarrow S^{6+} + 8e^{-}$** **reducing: $4(2e^{-} + O_2^{1-} \rightarrow 2O^{2-})$**
6. **3** H_2S + **2** HNO_3 → **3** S + **2** NO + **4** H_2O	**oxidizing: $3(S^{2-} \rightarrow S0 + 2e^{-})$** **reducing: $2(3e^{-} + N^{5+} \rightarrow N^{2+})$**
7. MnO_2 + $H_2C_2O_4$ + H_2SO_4 → $MnSO_4$ + **2** CO_2 + **2** H_2O	**oxidizing: $C_2^{6+} \rightarrow 2C^{4+} + 2e^{-}$** **reducing: $2e^{-} + Mn^{+4} \rightarrow Mn^{+2}$**
8. **2** H_2S + H_2SO_3 → **3** S + **3** H_2O	**oxidizing: $2(S^{2-} \rightarrow S^{0} + 2e^{-})$** **reducing: $2(2e^{-} + S^{0} \rightarrow S^{-2})$**
9. KIO_3 + **3** H_2SO_3 → KI + **3** H_2SO_4	**oxidizing $3(S^{4+} \rightarrow S^{6+} + 2e^{-})$** **reducing $6e^{-} + I^{5+} \rightarrow I^{-}$**
10. $K_2Cr_2O_7$ + **14** HCl → **2** KCl + **2** $CrCl_3$ + **3** Cl_2 + **7** H_2O	**oxidizing: $3(2Cl^{-} \rightarrow Cl_2^{0} + 2e^{-})$** **reducing: $6e^{-} + Cr_2^{6+} \rightarrow 2Cr^{3+}$**

Name________________

The Electrochemical Cell

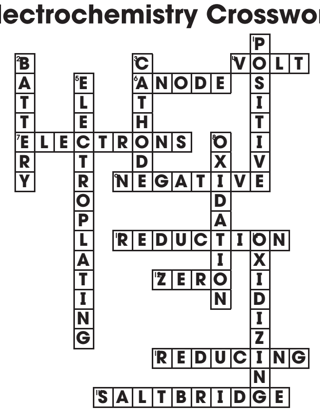


Answer each question, referring to the diagram and a table of standard electrode potentials.

1. Which is the more easily oxidized metal: aluminum or lead? **aluminum**
2. What is the balanced equation showing the spontaneous reaction that occurs? **$2Al^{0} + 3Pb^{+2} \rightarrow 2Al^{+3} + 3Pb^{0}$**
3. What is the maximum voltage that the above cell can produce? **+1.79 V**
4. What is the direction of electron flow in the wire? **from aluminum to lead**
5. What is the direction of positive ion flow in the salt bridge? **from aluminum nitrate to lead(II) nitrate**
6. Which electrode is decreasing in size? **aluminum**
7. Which electrode is increasing in size? **lead**
8. What is happening to the concentration of aluminum ions? **increasing**
9. What is happening to the concentration of lead ions? **decreasing**
10. What is the voltage in this cell when the reaction reaches equilibrium? **zero**
11. Which is the anode? **aluminum**
12. Which is the cathode? **lead**
13. Which is the positive electrode? **lead**
14. Which is the negative electrode? **aluminum**

Name________________

Electrochemistry Crossword



ACROSS

4. Unit of electrical potential
6. Electrode where oxidation takes place
7. Both atoms and ________ must be balanced in a redox equation.
9. The anode in an electrochemical cell has this charge.
10. Gain of electrons
12. Voltage of an electrochemical cell when it reaches equilibrium
13. A substance that is oxidized is the ________ agent.
14. Allows the flow of ions in an electrochemical cell

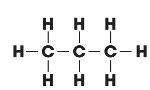
DOWN

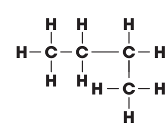
1. The anode in an electrolytic cell has this charge.
2. Another word for an electrochemical cell
3. Electrode where reduction takes place
5. Process of layering a metal onto a surface in an electrolytic cell
8. Loss of electrons
11. A substance that is reduced is the ________ agent.

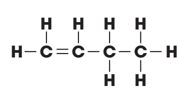
Name________________

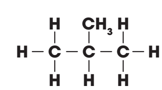
Naming Hydrocarbons

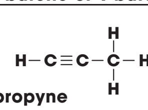
Name each compound according to the IUPAC naming system.

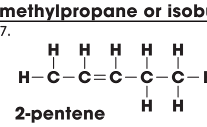
1.

propane

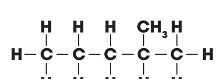
5.

butane

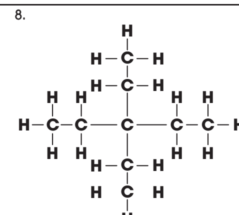
2.

n-butene or 1-butene

6.

methylpropane or isobutane

3.

propyne

7.

2-pentene

4.

2-methylpentane

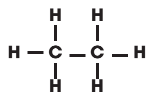
8.

3, 3-diethylpentane

Answer Key

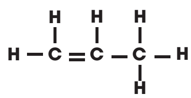
Name________________________

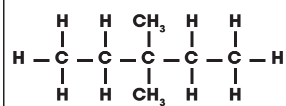
Structure of Hydrocarbons

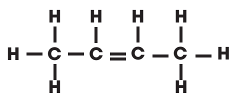
Draw the structure of each compound.

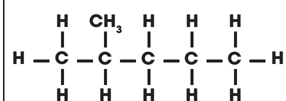
1. ethane


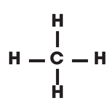
5. ethyne

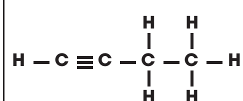

2. propene


6. 3, 3-dimethylpentane


3. 2-butene


7. 2, 3-dimethylpentane


4. methane


8. n-butyne


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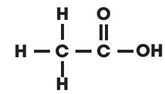
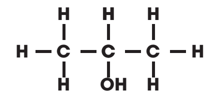
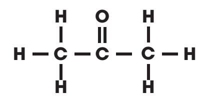
Name________________________

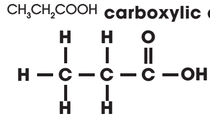
Functional Groups

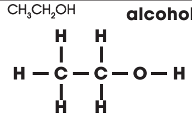
Classify each of the organic compounds as an *alcohol, carboxylic acid, aldehyde, ketone, ether,* or *ester*. Then, draw its structural formula.

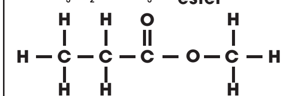
1. CH_3COOH **carboxylic acid**


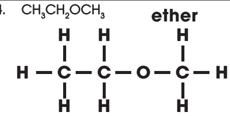
6. $CH_3CH(OH)CH_3$ **alcohol**


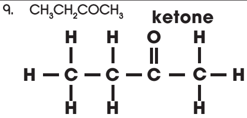
2. CH_3COCH_3 **ketone**


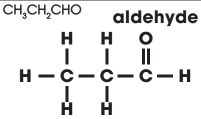
7. CH_3CH_2COOH **carboxylic acid**


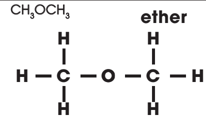
3. CH_3CH_2OH **alcohol**


8. $CH_3CH_2COOCH_3$ **ester**


4. $CH_3CH_2OCH_3$ **ether**


9. $CH_3CH_2COCH_3$ **ketone**


5. CH_3CH_2CHO **aldehyde**


10. CH_3OCH_3 **ether**


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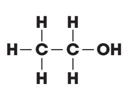
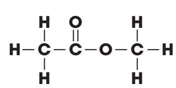
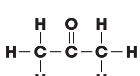
Name________________________

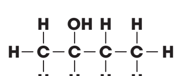
Naming Other Organic Compounds

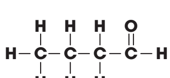
Name each compound.

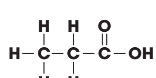
1.

ethanol

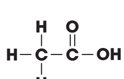
6.

methyl ethanoate or methyl acetate

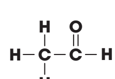
2.

propanone

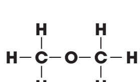
7.

2-butanol

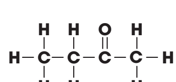
3.

butanol

8.

propanoic acid

4.

acetic acid or ethanoic acid

9.

ethanol

5.

dimethyl ether

10.

2-butanone

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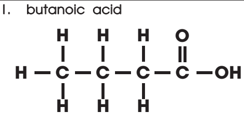
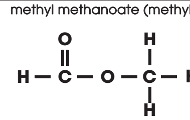
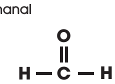
Name________________________

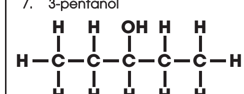
Structures of Other Organic Compounds

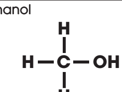
Draw the structure of each compound.

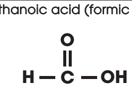
1. butanoic acid


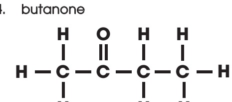
6. methyl methanoate (methyl formate)


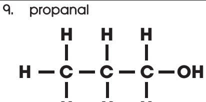
2. methanal


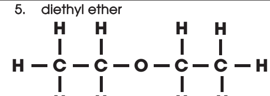
7. 3-pentanol


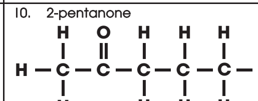
3. methanol


8. methanoic acid (formic acid)


4. butanone


9. propanal


5. diethyl ether


10. 2-pentanone


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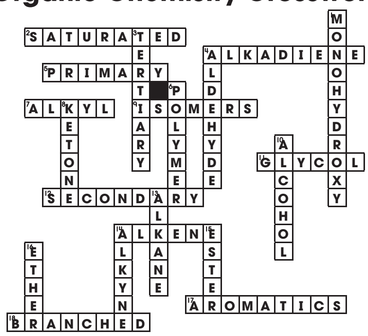
Answer Key

Name________________________

Organic Chemistry Crossword



Across

2. Hydrocarbons containing only single bonds
4. Contains two double bonds
5. Alcohol in which the hydroxyl group is attached to an end carbon
7. An alkane minus one hydrogen; It attaches to another carbon chain.
9. Compounds with the same molecular formula, but different structural formulas
11. A dihydroxy alcohol
12. Alcohol in which the hydroxyl is attached to a carbon attached to two other carbons
14. Open chain hydrocarbon containing one double bond
17. Organic compounds containing the benzene ring structure
18. Describes a hydrocarbon with a side chain of carbon atoms

Down

1. An alcohol with only one hydroxyl group in its structure
3. Alcohol in which the hydroxyl is attached to a carbon attached to three other carbons
4. General formula R-CHO
6. High molecular mass compound consisting of repeating units called monomers
8. General formula R-CO-R'
10. Contains one or more -OH groups
13. Saturated open chain hydrocarbon
14. Open chain hydrocarbon containing only one triple bond
15. Produced by the reaction of an alcohol and an acid
16. General formula R-O-R'

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