College-Prep Chemistry/Honors Chemistry

Unit 3: Classification & Separation of Matter

Unit Essential Question: In what ways can matter and the changes it undergoes be classified?

College-Prep: ~24 days, Honors: ~19 days

Topics

Phases of Matter

- 4 phases of matter (definitions, properties, examples)
- 6 phase-change processes
- phase-change heating and cooling curves
- phase diagrams (pressure vs. temperature)

Classification of Matter

- Pure Substances
    - Elements, Compounds
- Mixtures
    - Homogenous Mixtures
        - Colloids
        - Solutions
            - Solubility curves
            - Solution concentration (molarity, % mass, % volume)
    - Dilutions
    - Heterogeneous Mixtures
            - Suspensions

Chemical vs. Physical Changes

- Chemical and Physical Properties
- Definitions
- Distinguishing from Lab Observations
- Examples

Methods of Separation

- Know how to use these techniques, and when you would use them.
    - Filtration
    - Distillation
    - Crystallization
    - Chromatography: Paper and Gas
    - Centrifuge
    - Decanting
    - Charcoal Filtration (adsorption)
    - Electrolysis

Our virtues and our failings are inseparable, like force and ***matter****. When they* ***separate****, man is no more.*

Nikola Tesla
American Inventor
(1856-1943)

Supplements

- Mr. B's Google Site
- Student Slideshow (make a copy so you can edit)
- Teacher Slideshow
- Graphing the Phase Changes of Water lab
- Classifying Paper Clips activity
- Soda lab: Sugar Concentration
- ~~Physical/Chemical Change Lab~~
- Separation Techniques Lab
- Extension/Enrichment Options

Essential Questions

In what ways can matter and the changes it undergoes be classified?

Date	Essential Question	Summarize the Answer
	1) What are the 4 main phases of matter and the 6 main phase changes?	
	2) What happens to the temperature of a substance as it is changing phases?	
	3) What does the heating curve look like in going from a solid to a liquid to a gas?	
	4) **<u>WHY</u>** does the heating curve level off during the phase changes of melting and boiling?	
	5) What is a phase diagram and what does it communicate? x2	

Date	Essential Question	Summarize the Answer
	6) How are atoms and elements related?	
	7) What are the similarities and differences between metals, nonmetals, and metalloids?	
	8) What is the difference between a molecule and formula unit?	
	9) How do you determine whether a sample is an element, compound, or mixture?	
	10) What are 2 main types of mixtures and how are they different?	

<table>
<tr><th>Date</th><th>Essential Question</th><th>Summarize the Answer</th></tr>
<tr><td></td><td>11) How do you determine whether a sample is an element, compound, homogeneous mixture, or heterogeneous mixture?</td><td></td></tr>
<tr><td rowspan="2"></td><td>12) What are 3 types of solutions based on solubility?</td><td></td></tr>
<tr><td>13) What are 2 factors that affect the solubility of a solute?</td><td></td></tr>
<tr><td></td><td>14) What are 4 ways to measure solution concentration?</td><td></td></tr>
<tr><td></td><td>15) How do you properly make a solution?

x2</td><td></td></tr>
</table>

Date	Essential Question	Summarize the Answer
	16) How do you perform a serial dilution?	
	17) What is the relationship between sugar concentration and density?	
	18) How do you make a calibration curve?	
	19) How do you use a calibration curve to find an unknown variable?	
	20) How much sugar can be found in the average can (355 mL) of soda? Big Gulp (1183 mL)?	

Date	Essential Question	Summarize the Answer
	21) What is the difference between a physical change and a chemical change? x2	
	22) What methods can be used to separate mixtures and how do they work? x3	
	23) How do metric prefixes change the base unit?	

Vocabulary

CP Chemistry Quizlet Invite   Honors Chemistry Quizlet Invite

1. allotrope	
2. atom	
3. charcoal filtration	
4. chemical change	
5. chemical property	
6. chromatography	
7. colloid	
8. compound	
9. concentration (in reference to a solution)	
10. condensation	

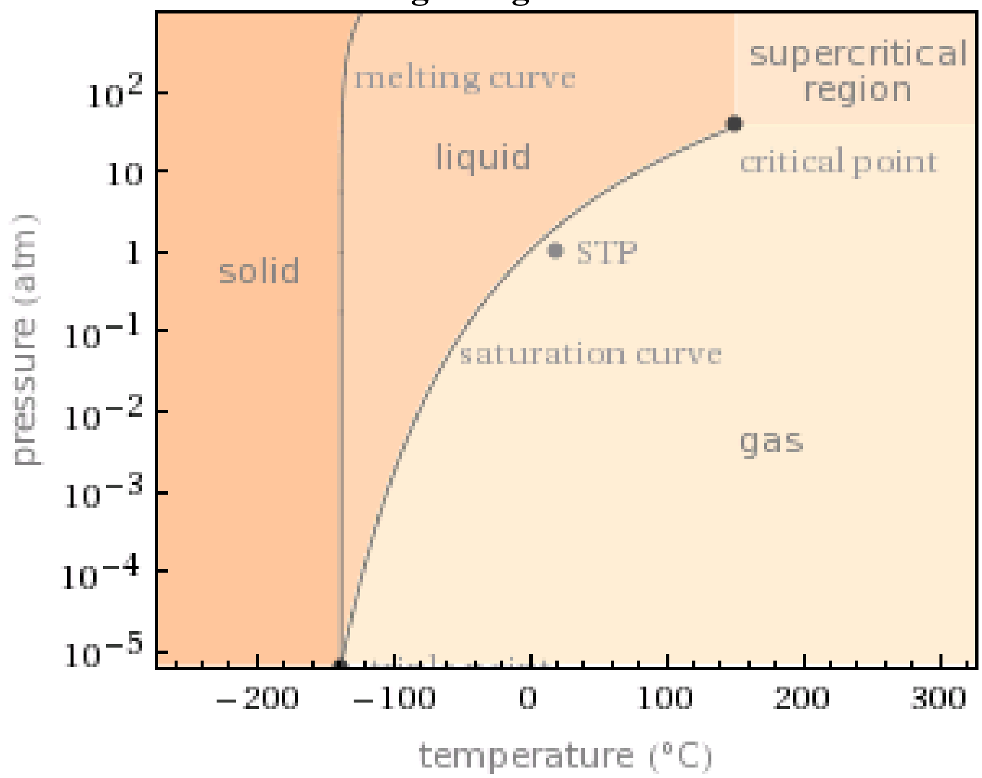
11. cooling curve	
12. crystallization	
13. decanting	
14. deposition	
15. distillation	
16. element	
17. evaporation	
18. filtration	
19. formula unit	
20. freezing	

21. gas	
22. heating curve	
23. heterogeneous mixture	
24. homogenous mixture	
25. isomer	
26. law of definite composition	
27. liquid	
28. melting	
29. mixture	
30. molarity	

31. molecule	
32. percent by mass (%m/m)	
33. percent by volume (%v/v)	
34. phase diagram	
35. physical change	
36. physical property	
37. plasma	
38. pure substance	
39. representative particle	
40. saturated solution	

41. solid	
42. solubility	
43. solution	
44. standard calibration curve	
45. standard solution	
46. sublimation	
47. supersaturated solution	
48. Tyndall effect	
49. unsaturated solution	

Phase Change Diagram of Butane



(Butane)

Analysis Questions: RTQ- <u>R</u>estate <u>T</u>he <u>Q</u>uestion as part of your Answer

1) At normal room conditions (1atm of pressure, and 25˚C), in which phase will butane exist?

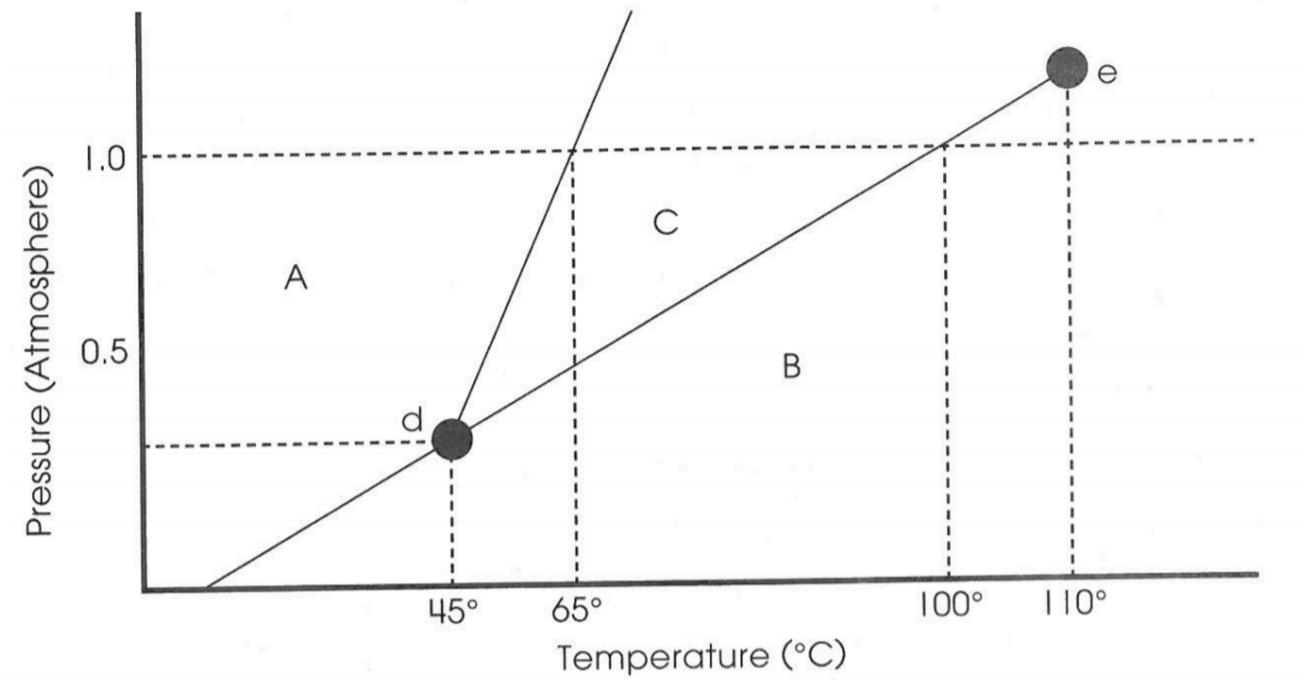
2) At normal pressure (1atm), at what temperature would butane freeze?

3) At normal pressure (1atm), at what temperature would butane boil?

4) A butane lighter contains liquid butane. For this butane to be a liquid at room temperature (25˚C), what would the minimum pressure need to be inside the lighter?

5) At what conditions could you get butane to go through the phase change of deposition?

Generic 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?

Phases of Matter Graphic Organizer

(Copy the image, insert a Google Drawing, paste the image into Google Drawing, and mark up the image.)

Indicate which phases of matter have the following properties. Also label the 6 different phase changes.

Mr.Balmer

Phases of Matter

Definition of matter:

Phase Changes

Properties Of Phases	Relatively high density: tightly packed particles	Relatively low density: loosely spread particles	Pressure doesn't greatly affect density: particles can't really be squeezed closer	Pressure greatly affects density: particles can be squeezed closer	Maintains its own shape in a container	Adopts the shape of its container	Expands to fill its container	Exists only at high temperatures and/or energies	Solid Example:	Liquid Example:	Gas Example:	Plasma Example:
solid phase	X											
liquid phase	X											
gas phase												
plasma phase												

Draw and **Label** a picture that represents each phase of matter or come up with a bumper sticker slogan for each phase of matter.

Phase Changes

Draw and Label a picture that represents each phase of matter or come up with a bumper sticker slogan for each phase of matter.

Solid

Liquid

Gas

Plasma

Gallery Walk: Periodic Table Samples (Honors Chem Skips This)

Metals are on the left side of the staircase. Nonmetals are on the right side of the staircase. Metalloids are on the staircase. Make-up version

Atomic #	Symbol	Name	Observations	Classification (metal, nonmetal, metalloid)
5	B			
11	Na			
12	Mg			
13	Al			
14	Si			
16	S			
19	K			
20	Ca			
24	Cr			
26	Fe			
27	Co			
28	Ni			
29	Cu			
30	Zn			
31	Ga			
34	Se			
35	Br			
48	Cd			
50	Sn			
51	Sb			
52	Te			
53	I			
78	Pt			
80	Hg			
82	Pb			
83	Bi			

Gallery Walk: Classification of Matter (Make-up version)

Sample #	Sample Name(s) and Formula(s)	ObservationsGallery Walk_types of matter_makeup	element, compound, homogenous mixture, or heterogeneous mixture,
1.			
2.			
3.			
4.			
5.			
6.			
7.			
8.			
9.			
10.			
11.			
12.			
13.			
14.			
15.			
16.			
17.			

Sample #	Sample Name(s) and Formula(s)	Observations	element, compound, homogenous mixture, or heterogeneous mixture,
18.			
19.			
20.			
21.			
22.			
23.			
24.			
25.	------------------	-------------------skip---------------	--
26.			
27.			
28.			
29.			
30.			
31.			
32.			
33.			
34.			
35.			

Graphic Organizer: Types of Matter

(Copy the image, insert a Google Drawing, paste the image into Google Drawing, and mark up the image.)

Key Terms: atoms, colloids, compounds, elements, formula units, heterogeneous mixtures, homogenous mixtures, mixtures, molecules, pure substances solutions, suspensions

Mr. Balmer

Types of Matter

____

- Unique set of physical and chemical properties
- Chemical symbol or formula can be use to identify them
- Obey the Law of Constant Proportions

____

- Matter made of one type of atom
- Found on the periodic table
- Each one has a unique number of protons (atomic number)

____

- Matter made of two or more **different** atoms chemically bonded
- Atoms can only be separated using a chemical change

____

- Two or more physically separable components
- Individual components retain their original properties
- Break the Law of Constant Proportions

____

- The components are **not** well-blended so the parts of the sample look **different**
- The properties throughout the sample **vary**

____

- The components are well-blended so all of the sample looks the **same**
- The properties throughout the sample are **uniform**

_____

- The smallest unit of matter that has the properties of the element
- Their nuclei are made out of neutrons and a unique number of protons.
- Their electrons are found in the orbitals (cloud or "rings).

_____

- Two or more atoms covalently bonded (sharing electrons)
- The smallest unit of matter that has the properties of a covalent compound (bonded nonmetals) or diatomic element (Br_2 I_2 N_2 Cl_2 H_2 O_2 F_2)

_____

- Two of more atoms ionically bonded (exchanging electrons)
- The smallest unit of matter that has the properties of an ionic compound (bonded metals and nonmetals)
- They dissociate (separate) into individual ions when dissolved.

_____

- Contain large, visible particles that tend to settle and be removed during filtration
- These samples are opaque (block light)

_____

- Contain suspended micro- to macroscopic particles
- Looks homo-genous to the naked eye
- Exhibit the Tyndall effect. You can see the scattered light beam in the sample.

_____

- Contain molecular or ionic-sized particles
- Does not exhibit the Tyndall effect. You cannot see the scattered light beam in the sample.

Colloids

Definition of Colloid:

	Dispersed Phase (minor component)			
		Gas	**Liquid**	**Solid**
Dispersion Medium (major component)	**Gas**	**None**	Liquid Aerosol	Solid Aerosol
	Liquid	Foam	Emulsion	Sol
	Solid	Solid Foam	Solid Emulsion (Gel)	Solid Sol

List an example of each and explain why you chose it.

Liquid Aerosol-

Solid Aerosol-

Foam-

Emulsion-

Sol-

Solid Foam-

Gel-

Solid Sol-

Element, Compound, or Mixture?

Scenario 1:

A colorless, homogeneous liquid is heated in a beaker. The liquid starts to boil and the amount of liquid slowly diminishes. After 20 minutes, the beaker is entirely empty and perfectly clean. Was the liquid in the beaker most likely an element, compound, or mixture? Explain your answer.

Scenario 2:

A substance consists of carbon, hydrogen, oxygen, and silicon. Is this substance most likely an element, compound, or mixture? Explain your answer.

Scenario 3:

A liquid with a yellowish appearance is poured through filter paper. The liquid that drips through ends up being colorless, while a yellow powder collects on the filter paper. Was the liquid in the beaker most likely an element, compound, or mixture? Explain your answer.

Scenario 4:
A 45.2 g sample of fool's gold is heated strongly and results in the formation of iron and sulfur. The fool's gold was not magnetic.

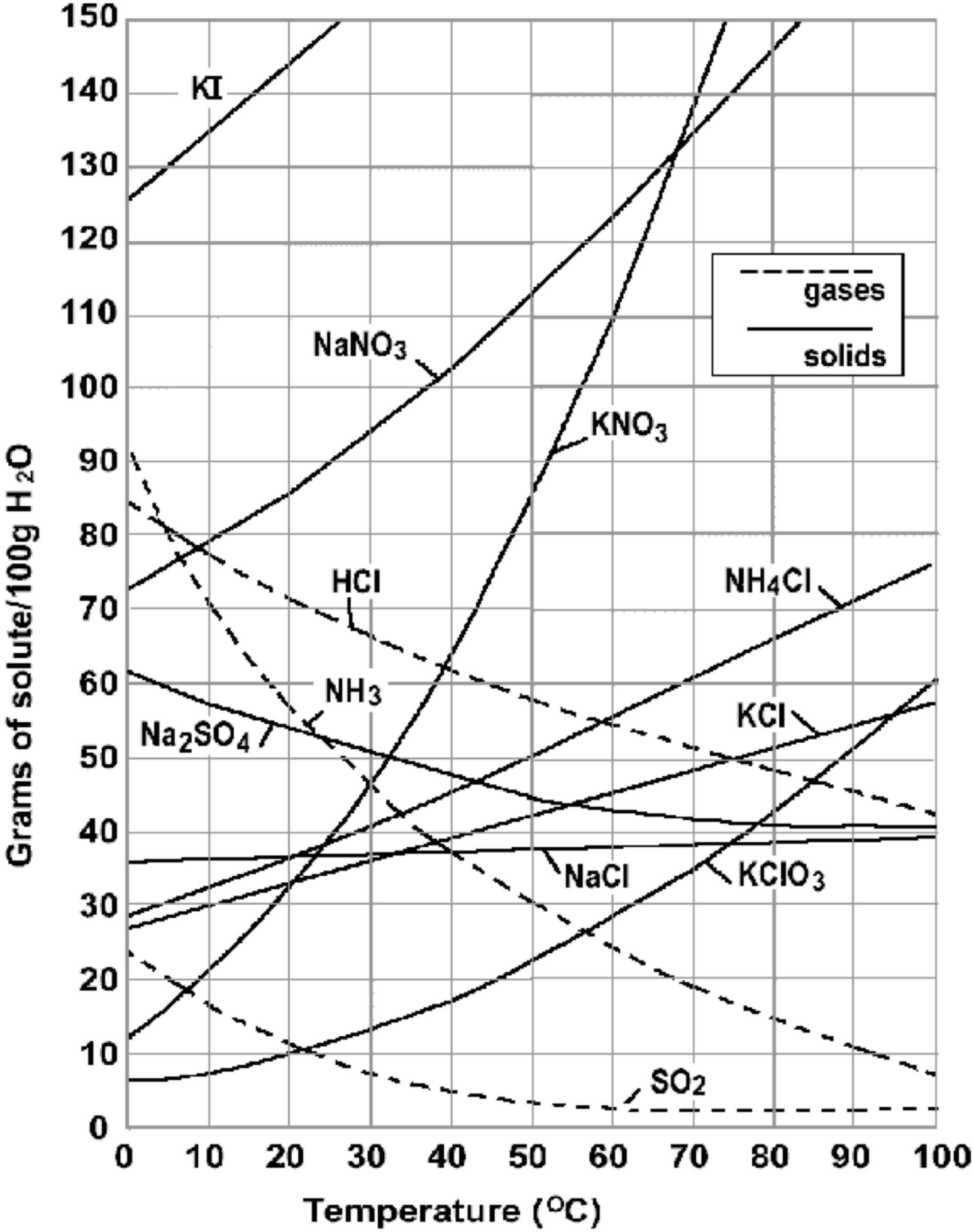
a) Is fool's gold an element, mixture or compound? Explain how you know.

b) If 21.0 grams of iron is produced, what mass of sulfur is produced? Explain.

Scenario 5:
A solid, metallic element is heated strongly in a flame. After a few seconds, the metal begins to glow as if on fire. The metal is pulled out of the flame, but continues to "glow" brightly for 15 more seconds. What once was the metal now has a white, powdery type of appearance. Is the white powder an element, compound, or mixture? Explain.

Scenario 6:
You walk back in a storeroom and find two beakers: one containing water, the other containing a mixture of hydrogen and oxygen. Compare and contrast the contents of the two beakers. (in what ways are they similar, in what ways are they different?)

Solubility Curve Graph



Solubility Curve Questions

Directions: Use the solubility curve on the previous page to answer the following questions.

1.) How do the solubilities of solids compare to the solubilities of gases?

2.) What is the solvent for the entire solubility curve?

3.) How much of that solvent does the entire graph refer to?

4.) What is the solubility of potassium chloride (KCl) at 90ºC?

5.) How much potassium nitrate (KNO_3) can 100 grams of water dissolve at 60ºC?

6.) What is the solubility of potassium iodide (KI) at 20ºC?

7.) What is the solubility of ammonia gas (NH_3) at 17ºC?

8.) How much ammonium chloride (NH_4Cl) can 200 grams of water dissolve at 90ºC?

9.) How much potassium chlorate ($KClO_3$) can dissolve in 500 grams of water at 70ºC?

10.) If 18 grams of sodium chloride (NaCl) is dissolved in 100 grams of water at 50ºC, how much more sodium chloride can be dissolved until the solution is saturated?

11.) If 70 grams of sodium sulfate (Na_2SO_4) are dissolved in 100 grams of water at 10ºC, is this solution unsaturated, saturated, or supersaturated?

12.) If 60 grams of sodium nitrate ($NaNO_3$) are dissolved in 100 grams of water at 0ºC, is this solution unsaturated, saturated, or supersaturated?

Answer the following questions based on the solubility chart below.

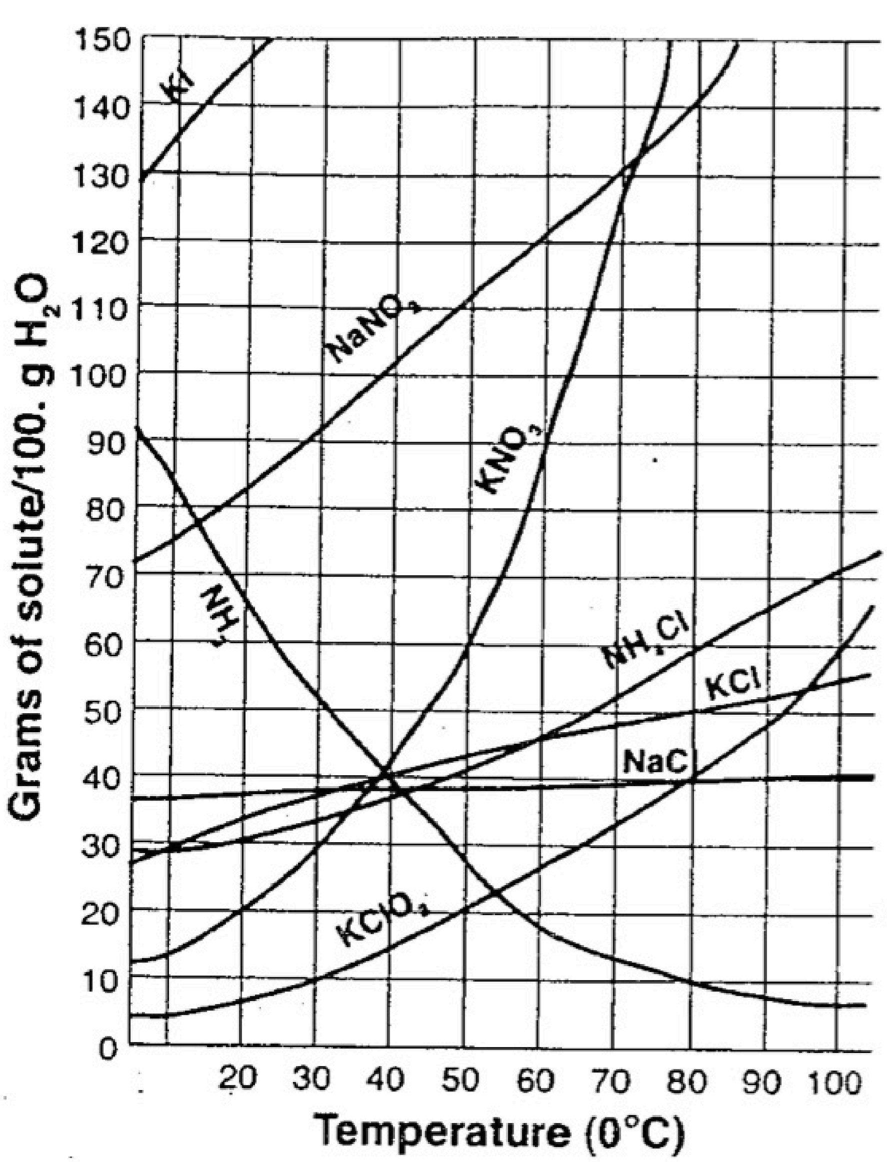
1) Which salt is least soluble in water at 20˚C?

2) How many grams of potassium chloride (KCl) can be dissolved in 200g of water at 80˚C?

3) At 40˚C, how much potassium nitrate (KNO_3) can be dissolved in 300g of water?

4) Which salt shows the least change in solubility from 0˚C to 100˚C?

5) At 30˚C, 90g of sodium nitrate ($NaNO_3$) is dissolved in 100g of water. Is this solution unsaturated, saturated, or supersaturated?



6) A saturated solution of potassium chlorate ($KClO_3$) is formed by dissolving enough potassium chlorate in 100g of water at 80˚C. If the solution is cooled to 50˚C, how many grams of potassium chlorate will precipitate?

7) Which 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?

Adapted from Instructional Fair ©

Solution Concentration Problems

Definitions

- Molarity (M) is the moles of solute divided by the volume of solution in liters. It has units of $\frac{\text{mol}}{\text{L}}$. This is the most common form of concentration that chemists use.

- % by mass (%m/m) is the mass of the solute divided by the total mass of the solution multiplied by one hundred. Since both masses have the same unit, they cancel, leaving you with no units.

- % by volume (%v/v) is the volume of the solute divided by the total volume of the solution multiplied by one hundred. Since both volumes have the same unit, they cancel, leaving you with no units. This is useful when both the solute and solvent are liquids.

Other Useful Measures of Concentration

- Molality (m) is the moles of solute divided by the mass of solvent in kilograms. Molality is especially useful with big changes in temperature. As the temperature of the solution changes, so does the volume. This would also affect the molarity. Molality would be unaffected by temperature changes because it is based on mass. It has units of $\frac{\text{mol}}{\text{kg}}$.
- Parts per million (ppm) is the amount of solute divided by the total amount of solution.
  - $\frac{\text{1g solute}}{10^6\text{g solution}}$ is 1ppm because there is 1g of solute and 1 million grams (Mega) of solution.
  - $\frac{\text{1mg solute}}{\text{1kg solution}}$ is also 1ppm because there are 1000mg in a gram, and there are 1000 grams in a kilogram, so there is $\frac{\text{1mg solute}}{10^6\text{mg solution}}$ or 1ppm.
- Parts per billion(ppb) and parts per trillion(ppt) work the same way.

Show all work.

1. A 4.00 g sugar cube (Sucrose: $C_{12}H_{22}O_{11}$) is dissolved in 350. mL of 80.0 °C water in a teacup. The density of water at 80 °C is 0.975 g/mL. One mole of sucrose weighs 342.34 g. One mole of water weighs 18.02 g. After adding the sugar cube, the solution has a volume of 351 mL
   a. What is the **molarity** of the sugar water?

   (0.0333 mol sugar/L solution)

   b. What is the **% by mass** of the sugar water?

   (1.16% m/m)

   c. What is the **molality** of the sugar water? (optional)

   (0.0342 mol sugar/kg water)

   d. What is the concentration of the sugar water in **parts per million**? (optional)

   (11,600 ppm)

2. Hydrochloric acid was once known as muriatic acid because it was the "marine" acid—it was made from seawater. Muriatic acid is still sold in many hardware stores and is used for cleaning bricks and tile. What is the **<u>molarity</u>** of this solution if 125 mL contains 27.3 grams of HCl? One mole of HCl has a mass of 36.46 g.

(5.99 mol HCl/L solution)

3. Ammonia, used in window cleaners, is relatively soluble in water. Calculate the **<u>molarity</u>** of a solution that contains 252 grams of NH_3 per liter. Ammonia has a molar mass of 17.04 g/mol.

(14.8 mol NH_3/L solution)

4. Rubbing alcohol (isopropyl alcohol) is often sold at a concentration of 70% by volume. The density of isopropyl alcohol is 0.785 g/mL. One mole of isopropyl alcohol weighs 60.10 g. The density of 70% isopropyl alcohol solution is 0.878 g/mL. The density of water is 0.997 g/mL. Assume that adding the volume of solute and the volume of solvent gives you the total volume of solution.
    a. What is the **molarity** of the rubbing alcohol?

    (9.14 mol alcohol/L solution)

    b. What is the **% by mass** of the rubbing alcohol?

    (62.6% m/m)

    c. What is the **molality** of the rubbing alcohol? (optional)

    (30.6 mol alcohol/kg water)

    d. What is the concentration of the alcohol in **parts per million**? (optional)

    (648,000 ppm)

5. The EPA limit for lead in drinking water is 15ppb or $\frac{15\text{g Pb}}{10^9\text{g solution}}$. The density of water at 25 °C is 0.997 g/mL. Since the solution is so dilute, you can assume that the density of the solution is the same as pure water. One mole of lead weighs 207.20 g. One mole of water weighs 18.02 g.

   a. What is the **molarity** of the lead in this water?

   (7.2×10^{-8} mol lead/L solution)

   b. What is the **% by mass** of the lead in this water?

   (1.5×10^{-6} %m/m)

   c. What is the **molality** of the lead in this water? (optional)

   (7.2×10^{-8} mol lead/kg water)

6. During a physical exam, a patient was found to have a cholesterol level of 1.60 mg per deciliter. If one mole of cholesterol has a mass of 386.67 g, what is the **molarity** of cholesterol in the patient's blood?

   (4.14×10^{-5} mol cholesterol/L blood)

7. People with diabetes have trouble regulating the sugar levels (glucose) in their bloodstream. Many people with diabetes take insulin shots to help regulate their sugar levels. If people get too much insulin, their sugar levels can drastically drop resulting in a condition known as hypoglycemia. This often occurs if blood glucose levels drop below 55 mg/dL. What is this concentration in **molarity**? Glucose has a molar mass of 180.18 g/mol.

   (3.1×10^{-3} mol glucose/L blood)

8. If people don't have enough insulin, their bodies can't convert glucose to glycogen. As a result, too much glucose builds up in the blood stream. This leads to a condition known as hyperglycemia. This often occurs if blood glucose levels rise above 180 mg/dL. What is this concentration in **molarity**? Glucose has a molar mass of 180.18 g/mol.

(1.0×10^{-2} mol glucose/L blood)

9. The EPA limit for cadmium in drinking water is 5 ppb or $\frac{5\text{g Cd}}{10^9\text{g solution}}$. What is this concentration in terms of **molarity**? Water has a density of 0.997 g/mL at 25 °C. You can assume the density of the solution is the same as pure water. Cadmium has a molar mass of 112.41 g/mol.

(4×10^{-8} mol cadmium/L solution)

Molality

10. Arsenic has a 50% lethal dose (LD_{50}) between 15 and 30 $mg_{As}/kg_{body\ weight}$. This means that 50% of the population receiving this dose would die. What is the concentration of these lethal doses in molality? Arsenic has a molar mass of 74.92 g/mol.

(2.0×10^{-4} mol arsenic/kg body weight;
4.0×10^{-4} mol arsenic/kg body weight)

How to Make a Solution

Modified from Flinn Scientific

Step 0: Calculate the quantity of solute needed for the desired concentration and solution volume.

Mass of solute (g) = Concentration (g/mL) * volume of solution (mL)

Solution Preparation Example: Make a 0.2000 g/mL solution of sugar water. We have 100.00 mL, 50.00 mL, and 25.00 mL volumetric flasks. To conserve resources, you are not allowed to use more than 8.00 g of sugar for any solution you make.

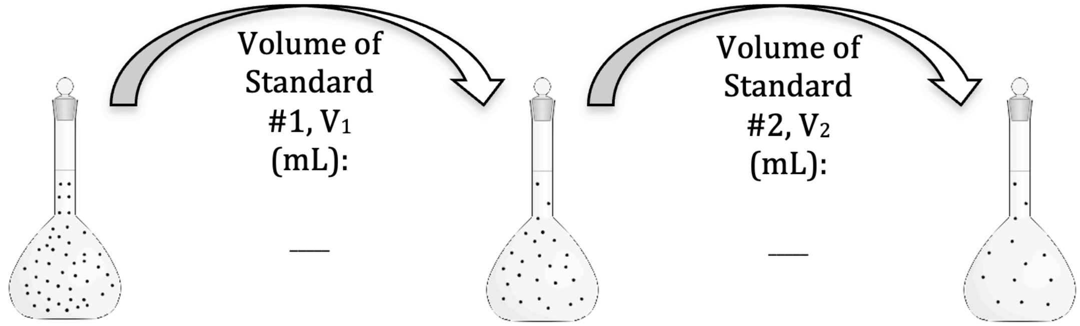
- How much sugar (in grams) would you need to make a 0.2000 g/mL solution in each of the volumetric flasks (100.00 mL, 50.00 mL, and 25.00 mL)?

- Which volumetric flask would you want to use? Why?

Step 1: Using a weighing dish, weigh the amount of solute you previously calculated in Step 0.	**Step 2:** Fill a volumetric flask ⅓ full with the appropriate solvent.	**Step 3:** Carefully transfer all of the solute from the weighing dish to a flask to dissolve. Rinse out weighing dish with a little extra solvent.	**Step 4:** Transfer solution to a volumetric flask. Add extra solvent or use a magnetic stir bar to dissolve the solute, if necessary.	**Step 5:** Add solvent up to the mark.	**Step 6:** Transfer your solution to a clean, labeled bottle.
					Chemical Name Solution Concentration (12M, 1g/mL, etc.) Date Name(s) Pd #

How to Dilute a Solution: Serial Dilution

Standard #1 **Standard #2** **Standard #3**



C_1 = 0.2000 g/mL

C_2 = 0.1200 g/mL
V_2 = 25.00 mL

C_3 = 0.05000 g/mL
V_3 = 25.00 mL

Serial Dilution Example:
Now that you have Standard #1, you will dilute some of it to make Standard #2. You will calculate how much of Standard #1 (V_1) you will need by using the equation below. The product of the first concentration and volume equals the product of the diluted concentration and volume.

$$C_1V_1 = C_2V_2 \quad \text{or} \quad V_1 = \frac{C_2V_2}{C_1}$$

•How much of Standard #1 (V_1) do you need to dilute to make 25.00 mL of Standard #2 with a concentration of 0.1200 g/mL?

Once you make Standard #2, you will calculate how much of it (V_2) you need to dilute to make Standard #3. You will use the same equation, but now you can change solutions 1 and 2 to solutions 2 and 3. The second equation is

$$C_2V_2 = C_3V_3 \text{ or } V_2 = \frac{C_3V_3}{C_2}$$

•How much of Standard #2 (V_2) do you need to dilute to make 25.00 mL of Standard #3 with a concentration of 0.0500 g/mL?

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 x V, then $M_1 \times V_1 = M_2 \times V_2$. This dilution equation works for all dilutions, not just those of an acid. It also works for other units of concentration besides molarity like g/mL, %v/v, or %m/m. For that reason, the equation is sometimes written $C_1 \times V_1 = C_2 \times V_2$, where C stands for concentration. For this worksheet, you'll need to assume volumes are additive (25 mL + 25 mL = 50 mL). This isn't always the best assumption, though.

Solve each problem and show all work. See my teacher slideshow for more detailed answers.

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

83 mL of 18 M sulfuric acid

2) How much concentrated 12 M hydrochloric acid is needed to prepare 100 mL of a 2.0 M solution?

17 mL of 12 M HCl

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

125 mL of 3.0 M HNO_3

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

~100 mL H_2O (V_2 = 150 mL total solution)

5) How much water should be added to 100. mL of 18 M sulfuric acid to prepare a 1.5 M solution?

~1.1 L of H_2O (V_2 = 1.2 L or 1200 mL total solution)

Safety Note: Be careful when diluting acids and bases. They often generate lots of heat- enough to crack the glassware they are in. Always add acid to water and use an ice bath if required.

Dilution Worksheet

$C_1V_1 = C_2V_2$
or
$M_1V_1 = M_2V_2$
For this worksheet, you'll need to assume volumes are additive (25 mL + 25 mL = 50 mL).
This isn't always the best assumption, though.

1) If I add 25 mL of water to 125 mL of a 0.15 M NaOH solution, what will the molarity of the diluted solution be?

M_2 = 0.125 M

2) If I add water to 100 mL of a 0.15 M NaOH solution until the final volume is 150 mL, what will the molarity of the diluted solution be?

M_2 = 0.100 M

3) How much 0.05 M HCl solution can be made by diluting 250 mL of 10M HCl?

V_2 = 50,000 mL

4) I have 345 mL of a 1.5 M NaCl solution. If I boil the water until the volume of the solution is 250 mL, what will the molarity of the solution be?

M_2 = 2.07 M

5) How much water would I need to add to 500 mL of a 2.4 M KCl solution to make a 1.0 M solution?

V_2 = 1200 mL so ~700mL of water

POGILY: Physical and Chemical Changes

Process Oriented Guided Inquiry Learning, Yay!

Target Inquiry GVSU-2009, Chad Brindle, Grandville High School ’14-‘15

Predictions:

1) Take a moment and think about what the difference between the terms *physical change* and *chemical change* might be in the world of chemistry. Write down how you would describe these two chemistry terms.

2) Now, share your ideas with your lab partner(s). How are your ideas the same and how are they different?

3) Once you’ve compared your ideas, look up the definitions for physical changes and chemical changes. How were you definitions similar or different?

Materials: One set of *Physical and Chemical Changes* cards per lab group (9 cards each, Lettered A-I)

Procedure: You have been provided with nine *Physical and Chemical Changes* cards. Each card represents a change that a substance or substances undergo.

The circle(s) on the left represent the atoms and molecules before they undergo the change and the circle(s)on the right represent the atoms and molecules after they undergo the change. Each circle is a snapshot of a situation, so the number of particles between “before” and “after” may not be the same. Particles that have a + or – are called ions. Something about that particle has changed and the “charged” particle has different properties from the “neutral” particle.

Your group’s task is to categorize each of these changes as “physical” or “chemical.” This means you must decide, based on your observations of the changes, what indicates a physical change has occurred and what indicates a chemical change has occurred.

As your group works through categorizing the changes, record your ideas in the *Data and Observations* section. First, describe what seems to be occurring during the change. Once you’ve decided whether a change is physical or chemical, record your decision. Then, in the space provided, describe why your group decided a change was physical or chemical.

Data and Observations: Data Cards

Change A Describe the change: Type of Change (physical or chemical): ________________ Defend your decision:	**Change F** Describe the change: Type of Change (physical or chemical): ________________ Defend your decision:
Change B Describe the change: Type of Change (physical or chemical): ________________ Defend your decision:	**Change G** Describe the change: Type of Change (physical or chemical): ________________ Defend your decision:
Change C Describe the change: Type of Change (physical or chemical): ________________ Defend your decision:	**Change H** Describe the change: Type of Change (physical or chemical): ________________ Defend your decision:
Change D Describe the change: Type of Change (physical or chemical): ________________ Defend your decision:	**Change I** Describe the change: Type of Change (physical or chemical): ________________ Defend your decision:
Change E Describe the change: Type of Change (physical or chemical): ________________ Defend your decision:	

Results and Analysis: Based on your lab group's discussions and decisions, write concise, one-sentence descriptions for *physical change* and *chemical change*.

1. Physical change:

2. Chemical Change:

Connections to the Real World: Each of the changes you observed represents a real process that you probably have seen or heard about. Obtain an *Atom Key* from your teacher to help you identify what processes are occurring in each of the *Physical and Chemical Changes* cards. Below are descriptions of each of these processes in some real-world situation. Match each *Physical and Chemical Changes* card with the appropriate description.

3. ______ Road Salt, sodium chloride, is commonly used to de-ice roads during the winter. When road salt dissolves in the water on the road, it reduces the temperature at which the water would freeze. This helps prevent ice from forming on the roads.

4. ______ Cars have been made from steel, which is mostly iron, since their introduction into society in the early 20th century. When the iron interacts with oxygen, it forms rust. This problem is accelerated by the wet, salty roads in many cold, winter climates.

5. ______ Copper has been used by humans for about 10,000 years. Due to its excellent flexibility and great ability to conduct electricity, copper is used for electrical wires as well as in pipes for plumbing.

6. ______ Original camera "flash bulbs" consisted of very fine magnesium filaments. An electrical current, triggered by the camera shutter, heats the filament until it ignites and burns, very quickly and brightly, with the oxygen in the air.

7. ______ Before the invention of the refrigerator, perishable food was often kept in ice boxes, which were cooled using blocks of ice. People relied on ice boxes even during the summer months. Ice was often stockpiled in large "ice houses" during the winter and could often be kept from melting until the following winter.

8. ______ Henry Cavendish is credited with identifying hydrogen gas as a unique element in 1766. Cavendish produced hydrogen gas by combining a metal, such as magnesium, with a strong acid, such as hydrochloric acid. Hydrogen production soon became useful as balloonists found this "lighter than air" gas quite useful.

9. ______ Sodium bicarbonate, commonly known as baking soda, is used to make baked goods light and fluffy. It does so by decomposing to produce water vapor and carbon dioxide gas. This decomposition process is initiated by the presence of an acid, which donates hydrogen ions.

10. ______ The first train locomotives were powered by steam. A very hot fire, usually coal powered, heats a large tank of water called a boiler. When the water evaporates and turns to steam, the boiler becomes highly pressurized by the steam. The high pressure steam is then used to push large pistons, which turns the wheels.

11. ______ Rubbing alcohol has many uses, most commonly as an antiseptic for cleaning minor cuts or contaminated surfaces. Household rubbing alcohol is a mixture of isopropanol and water.

Questions: Consider the following two situations. After reading the description, decide whether the change described is physical or chemical, according to the definitions you created. Then, in the space provided, illustrate the change at the particulate level. You can insert a Google Drawing to make your pictures. Use the same colors from the *Atom Key* in order to correctly represent the particle(s).

12. Pure copper can be heated until it melts just below 2000 °F. The liquid copper can be poured into molds.

Physical or Chemical Change:

Before

After

13. Natural gas consists mostly of methane, CH_4. Natural gas is used to heat homes and cook food. The methane is "burned" by reacting with oxygen, O_2. The products of this reaction are carbon dioxide, CO_2 and water, H_2O.

Physical or Chemical Change:

Before

After

Physical Vs. Chemical Properties

A physical property can be observed or measured without changing the chemical composition of the material. The substance you started with at the beginning will still be the same substance at the end. The chemical symbol/formula for the substance will not change. For example, color, shape, mass, length, and odor are all examples of physical properties.

A chemical property indicates how a substance will react with another substance. It can only be observed or measured when the substance is reacting with another substance to form a different chemical. The original substance will no longer be the same substance. 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 (III) oxide, which is a different substance.

Classify the following properties as either chemical or physical properties by putting a check in the appropriate column.

Property	**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		

Physical Vs. Chemical Changes

In a physical change, a substance changes, but its chemical composition stays the same. The substance before and after the change is the same substance so it has the same chemical symbol/formula.

In a chemical change, a new substance is produced. The original chemical symbol(s)/formula(s) is not the same as the ending chemical symbol(s)/formula(s).

Classify the following changes as physical changes (P) or chemical changes (C).

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. ____ Evaporation

9. ____ Ice melting

10. ____ Milk that has gone "sour" after its expiration date

11. ____ Sugar dissolves in water.

12. ____ Rotting wood turning back into forest nutrient

13. ____ Pancakes cooking on a griddle

14. ____ Grass 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.