Table T — Concentration

$$\text{parts per million} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 1000000$$

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

Overview:

The expressions given are used to indicate the concentration or strength of a solution. A solution consists of two components: the solute, or dissolved substance, and the solvent, or the dissolving substance. The most common solvent is water. In the parts per million equation, the grams of solution is the total mass of the solute and the solvent.

Example – parts per million:

An aqueous solution has 0.04 gram of oxygen dissolved in 1000. grams of water. Calculate the dissolved oxygen concentration of this solution in parts per million.

Equation: $\text{parts per million} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 1\,000\,000$

Solution: grams of solute = 0.04 gram

grams of solution = 1000. grams of water + 0.04 grams of oxygen = 1000.04 grams

Substitution: $\text{parts per million} = \frac{0.04 \text{ gram}}{1000.04 \text{ grams}} \times 1\,000\,000$

Answer: Concentration = 40. ppm

Example – molarity:

What is the molarity of a solution that contains 0.50 mole of NaOH in 0.50 liter of solution?

Equation: $\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$

Substitution: $\text{molarity} = \frac{0.50 \text{ mole of NaOH}}{0.50 \text{ liter of solution}}$

Answer: Concentration = 1.0 M

Additional Information:

- If the molarity of a solution is known, the number of moles of solute in a given volume of that solution can be found using the equation: moles of solute = molarity × liters of solution.

Set 1 — Concentration – ppm

1. What is the concentration of a solution, in parts per million, if 0.02 gram of Na_3PO_4 is dissolved in 1000 grams of water?

   (1) 20 ppm (3) 0.2 ppm
   (2) 2 ppm (4) 0.02 ppm 1 ____

2. What is the concentration of a solution in parts per million, if 0.089 gram of NaCl is dissolved in 250 grams of water?

   (1) 0.03 ppm (3) 35.6 ppm
   (2) 0.27 ppm (4) 356 ppm 2 ____

3. An aqueous solution has 0.0070 gram of oxygen dissolved in 1000. grams of water. Calculate the dissolved oxygen concentration of this solution in parts per million. Your response must include both a correct numerical setup and the calculated result.

Base your answer to question 4 using the information below and your knowledge of chemistry.

A town located downstream from a chemical plant was concerned about fluoride ions from the plant leaking into its drinking water. According to the Environmental Protection Agency, the fluoride ion concentration in drinking water cannot exceed 4 ppm. The town hired a chemist to analyze its water. The chemist determined that a 175-gram sample of the town's water contains 0.000250 gram of fluoride ions.

4. How many parts per million of fluoride ions are present in the analyzed sample?

5. More than 60 parts per million of dissolved minerals is considered to be hard water. A 750-gram sample of water was found to contain 0.04 gram of minerals. Would this sample of water be considered hard water? Explain your answer.

____

____

Set 1 — Concentration – molarity

.. Molarity is defined as the

(1) moles of solute per kilogram of solvent
(2) moles of solute per liter of solution
(3) mass of a solution
(4) volume of a solvent 1 _____

2. A 3.0 M HCl(aq) solution contains a total of

(1) 3.0 grams of HCl per liter of water
(2) 3.0 grams of HCl per mole of solution
(3) 3.0 moles of HCl per liter of solution
(4) 3.0 moles of HCl per mole of water

2 _____

3. What is the molarity of a solution that contains 0.50 mole of NaOH in 0.50 liter of solution?

(1) 1.0 M (3) 0.25 M
(2) 2.0 M (4) 0.50 M 3 _____

4. What is the total number of moles of NaCl(s) needed to make 3.0 liters of a 2.0 M NaCl solution?

(1) 1.0 mol (3) 6.0 mol
(2) 0.70 mol (4) 8.0 mol 4 _____

5. What is the total number of moles of solute in 4.0 liters of 4.0 M NaOH?

(1) 1.0 mole (3) 3.0 moles
(2) 2.0 moles (4) 16 moles 5 _____

Base your answers to question 6 on the information below.

A student is instructed to make 0.250 liter of a 0.200 M aqueous solution of $Ca(NO_3)_2$.

6. *a)* In the space below, show a correct numerical setup for calculating the total number of moles of $Ca(NO_3)_2$ needed to make 0.250 liter of the 0.200 M calcium nitrate solution.

b) In order to prepare the described solution in the laboratory, two quantities must be measured accurately. One of these quantities is the volume of the solution. What other quantity must be measured to prepare this solution?

CONCENTRATION

Show work for each of the following problems:

1. A 135 g sample of seawater is evaporated to dryness, leaving 4.73 g of salts. Calculate the mass percent of solute present in the seawater.

2. Cow's milk typically contains 4.5% by mass of lactose, $C_{12}H_{22}O_{11}$. Calculate the mass of lactose present in 175 g of milk.

3. Calculate the molarity of a solution prepared by dissolving 1.56 g of gaseous HCl into enough water to make 26.8 mL of solution.

4. Calculate the molarity of a solution prepared by dissolving 11.5 g of solid NaOH in enough water to make 1.50 L f solution.

5. How much solid $K_2Cr_2O_7$ (molar mass = 294.2 g) must be weighed out to make 2.00 L of an aqueous 0.200 M solution?

6. In the US, drinking water cannot contain more than 5×10^{-4} mg of mercury per gram of sample. What would that be in parts per million (ppm)?

7. The Dead Sea contains 4.64 kg of bromide ion in 1000. kg of water. Calculate the bromide ion concentration in ppm.

8. Household chlorine bleach contains 5.00% by mass of sodium hypochlorite, NaClO. In a 500 g sample, what mass is NaClO?

9. What is the volume of a 1.5 M NaCl solution made with 174 g of salt?

10. A tincture (alcohol solution) of iodine is prepared by dissolving 5.15 g of I_2 in enough alcohol to make 225 mL of solution. What is the molarity?

Solutions

Q1 Answer these questions about solubility:

a) Give a definition of **solubility**.

b) What factor can affect the solubility of a substance?

Q2 What's the **molarity** of a salt solution with 0.5 moles of salt dissolved in water to form 250 mL of solution?

Hint: work out the molar mass of the compounds first — then you can see how many moles you've got. See Section 2 if you've forgotten about moles.

Q3 Find the **molarity** of the following solutions:

a) 40 g of sodium hydroxide (NaOH) dissolved in water to form 500 mL of solution.

b) 202 g of potassium nitrate (KNO_3) dissolved in water to form 1500.0 mL of solution.

c) 36.5 g of hydrogen chloride gas dissolved in water to form 1000 mL of solution.

Q4 Find the **percent by volume** concentration of the following solutions:

a) 25 mL of LiOH dissolved in water to make 3.0 liters of solution.

b) 4.65 mL of $CuCl_2$ dissolved to make 35.0 mL of solution.

c) 50.00 mL of a 50% (by volume) NaCl solution added to more solvent to make 120.00 mL of solution.

Q5 Find the **percent by mass** concentration of the following solutions:

a) 800.0 g of solution containing 23.00 g of NaOH.

b) 7.500 g of $SrCl_2$ dissolved to form 178.0 g of solution.

Q6 Find the **parts per million** concentration of the following solutions:

a) 78 g of MgO dissolved to form 2.0 kg of solution.

b) 675 g of KI solution containing 90.0 g of KI.

Q7 Using the *Reference Tables*, decide whether each of the following solutions is **unsaturated**, **saturated**, or **supersaturated**.

a) 100 g of $NaNO_3$ dissolved in 100 g of water at 40 °C

b) 10 g of NH_4Cl dissolved in 20 g of water at 40 °C

c) 200 g of NaCl dissolved in 500 g of water at 95 °C

d) 100 g of KCl dissolved in 1,000 g of water at 25 °C

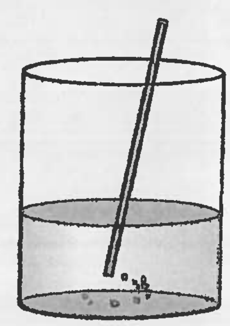
Q8 10 g of a nonvolatile solute are added to a beaker of water.

a) Explain what is meant by a **nonvolatile solute**.

b) What effect will this have on the **melting** and **boiling point** of the water?

The boiling point was changed by 2.5 °C by the addition of the solute.

c) If an extra 10 g of solute was added, what would happen to the boiling point?



Molarity Practice

$$M = \frac{\text{Moles}}{\text{Liters}}$$

A. Find the moles in:

1. 250 mL of a 4M soln
2. 600 mL of a .2M soln
3. 2L of a 5M soln
4. .5L of a .03M soln

B. Find the Molarity of:

5. 4 moles of solute in 100 mL of soln
6. .5 moles of HCl in 2 L of soln
7. 0.1 mole in 1000 mL of solution
8. 12.8 moles of KCl in 3.2 L of soln

C. Find the grams in

9. 6L of .3M NH_3 solution
10. 0.250 L of 5M HNO_3 (aq)
11. 1 L of .2M HCl (aq)
12. 1.50 L of 1 M $CuSO_4$ solution

1. How many grams of ammonium chloride (gram formula mass = 53.5 g) are contained in 0.500 L of a 2.00 M solution?
   (1) 10.0 g (3) 53.5 g
   (2) 26.5 g (4) 107 g

2. What is the total number of moles of solute contained in 0.50 liter of 3.0 M HCl?
   (1) 1.0 (3) 3.0
   (2) 1.5 (4) 3.5

3. What is the molarity of a solution containing 20. grams of NaOH in 0.50 liter of solution?
   (1) 1.0 (3) 0.50
   (2) 2.0 (4) 10.

4. The molarity (M) of a solution is equal to the
   (1) $\frac{\text{number of grams of solute}}{\text{liter of solvent}}$
   (2) $\frac{\text{number of grams of solute}}{\text{liter of solution}}$
   (3) $\frac{\text{number of moles of solute}}{\text{liter of solvent}}$
   (4) $\frac{\text{number of moles of solute}}{\text{liter of solution}}$

5. What is the concentration of a solution which contains 1 mole of $CaCl_2$ dissolved in 2,000 milliliters of solution?
   (1) 1 M (3) 0.5 M
   (2) 2 M (4) 0.25 M

6. What is the molarity of an H_2SO_4 solution if 0.25 liter of the solution contains 0.75 mole of H_2SO_4?
   (1) 0.33 M (3) 3.0 M
   (2) 0.75 M (4) 6.0 M

7. What is the molarity of a solution that contains 4 grams of NaOH in 500 milliliters of solution? [Formula mass of NaOH = 40.]
   (1) 0.1 M (3) 0.2 M
   (2) 2 M (4) 0.5 M

8. How many moles of KNO_3 are required to make 0.50 liter of a 2.0 M solution of KNO_3?
   (1) 1.0 (3) 0.50
   (2) 2.0 (4) 4.0

9. How many grams of KOH are needed to prepare 250. milliliters of a 2.00 M solution of KOH (formula mass = 56.0)?
   (1) 1.00 g (3) 28.0 g
   (2) 2.00 g (4) 112 g

10. How many milliliters of 12.0 M HCl(aq) must be diluted with water to make exactly 500. mL of 3.00 M hydrochloric acid?
   (1) 100. mL (3) 200. mL
   (2) 125. mL (4) 250. mL

11. What is the molarity of a solution that contains 40. grams of NaOH in 0.50 liter of solution?
   (1) 1.0 M (3) 0.50 M
   (2) 2.0 M (4) 0.25 M

12. What is the total number of moles of solute in 2.0 liters of 3.0 M NaOH?
   (1) 1.0 mole (3) 3.0 moles
   (2) 2.0 moles (4) 6.0 moles

Name ______________________________ Date ______________

Concentration Questions

____1. The number of Joules needed to raise the temperature of 10 grams of water from 20°C to 30°C is
1) 42 3) 420
2) 84 4) 1680

____2. What is the gram formula mass of K_2CO_3?
1) 138 g 3) 99 g
2) 106 g 4) 67 g

____3. Which quantity is equivalent to 39 grams of LiF?
1) 1.0 mole 3) 0.50 mole
2) 2.0 moles 4) 1.5 moles

____4. Which solution is most concentrated?
1) 0.1 mole of solute dissolved in 400 ml of solvent
2) 0.2 mole of solute dissolved in 300 ml of solvent
3) 0.3 mole of solute dissolved in 200 ml of solvent
4) 0.4 mole of solute dissolved in 100 ml of solvent

____5. If 0.025 gram of $Pb(NO_3)_2$ is dissolved in 100. grams of H_2O, what is the concentration of the resulting solution, in parts per million?
1) 2.5×10^{-4} ppm 3) 250 ppm
2) 2.5 ppm 4) 4.0×10^{3} ppm

____6. How many grams of KNO_3 should be dissolved in water to make 500.0 grams of a 20.0 ppm solution?
1) 1.00×10^{-1} g 3) 1.00×10^{-3} g
2) 1.00×10^{-2} g 4) 1.00×10^{-4} g

7. Show a correct numerical setup for determining how many liters of a 1.2 M solution can be prepared with 0.50 mole of $C_6H_{12}O_6$.

8. Base your answer to the following question on the information below.

A student is instructed to make 0.250 liter of a 0.200 M aqueous solution of $Ca(NO_3)_2$.

Show a correct numerical setup for calculating the total number of moles of $Ca(NO_3)_2$ needed to make 0.250 liter of the 0.200 M calcium nitrate solution.

9. An aqueous solution contains 300. parts per million of KOH. Determine the number of grams of KOH present in 1000. grams of this solution.

10. Base your answer to the following question on the information below.

The health of fish depends on the amount of oxygen dissolved in the water. A dissolved oxygen (DO) concentration between 6 parts per million and 8 parts per million is best for fish health. A DO concentration greater than 1 part per million is necessary for fish survival.

Fish health is also affected by water temperature and concentrations of dissolved ammonia, hydrogen sulfide, chloride compounds, and nitrate compounds. Most freshwater fish thrive in water with a pH between 6.5 and 8.5.

A student's fish tank contains fish, green plants, and 3800 grams of fish-tank water with 2.7 x 10-2 gram of dissolved oxygen. Phenolphthalein tests colorless and bromthymol blue tests blue in samples of the fish-tank water.

Determine if the DO concentration in the fish tank is healthy for fish. Your response must include:

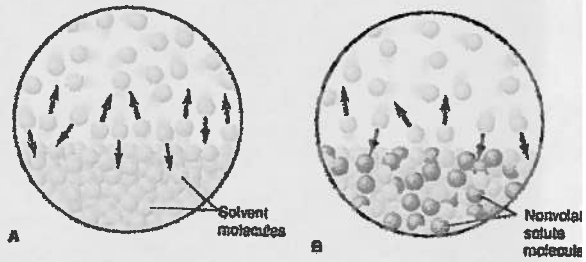
- a correct numerical setup to calculate the DO concentration in the water in parts per million
- the calculated result
- a statement using your calculated result that tells why the DO concentration in the water is or is not healthy for fish

Topic 10: Solutions

10.5: Colligative Properties

AIM: ____

- ***Colligative Properties***



Why do we put salt on the sidewalks when it snows?

- When nonvolatile ______________________ ______________________
- When salt ______________________ ______________________
    - Adding salt ______________________ ______________________
    - 1 mole of ______________________
        - Ex: 1 mole of ______________________

- ***Molecular vs. Ionic***

What happens when we put a molecular (covalent) substance into water?

- When ______________________

What happens when we add an ionic substance into water?

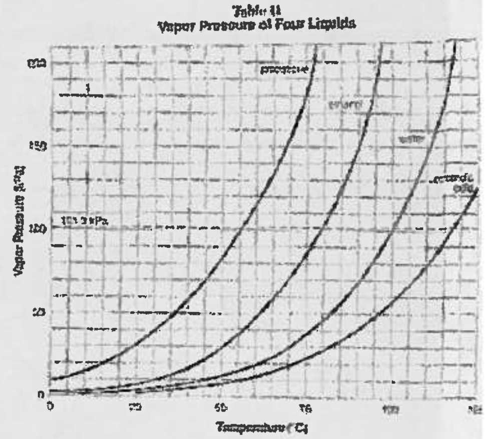
- When ______________________

- 1 mole ______________________
    - Will ______________________
    - The greater ______________________
        - $CaCl_2$ = ______________________

What happens to the boiling point when you add something to water?

- One ______________________
    - 1 mole of ______________________
    - 1 mole of ______________________

- ***Vapor Pressure***
    - The pressure that a ______________________________
    - Table H Reference Tables
        - 4 substances
        - Propanone ______________________________
          ____
        - Ethanoic acid ____
      ____



- ***Boiling Point***
    - As temperature ____
        - When Vapor pressure = ____
            - Liquid becomes a ______________________________
            - When bubbles ____
        - Normal boiling point____
          ____
    - Heat of Vaporization ____
      ____
    - Normal boiling point of water = ____
    - Vapor pressure:____
        - Table H = ____
          ____
            - When the pressure ____
                - If the pressure is ____
                  ____

Regents Questions

- The heat energy to change one mole of liquid into a vapor at the boiling point is called the

(1) Heat of vaporization
(2) Heat of formation
(3) Heat of solution
(4) Heat of fusion

- Water boils at 90°C when the pressure exerted on the liquid equals

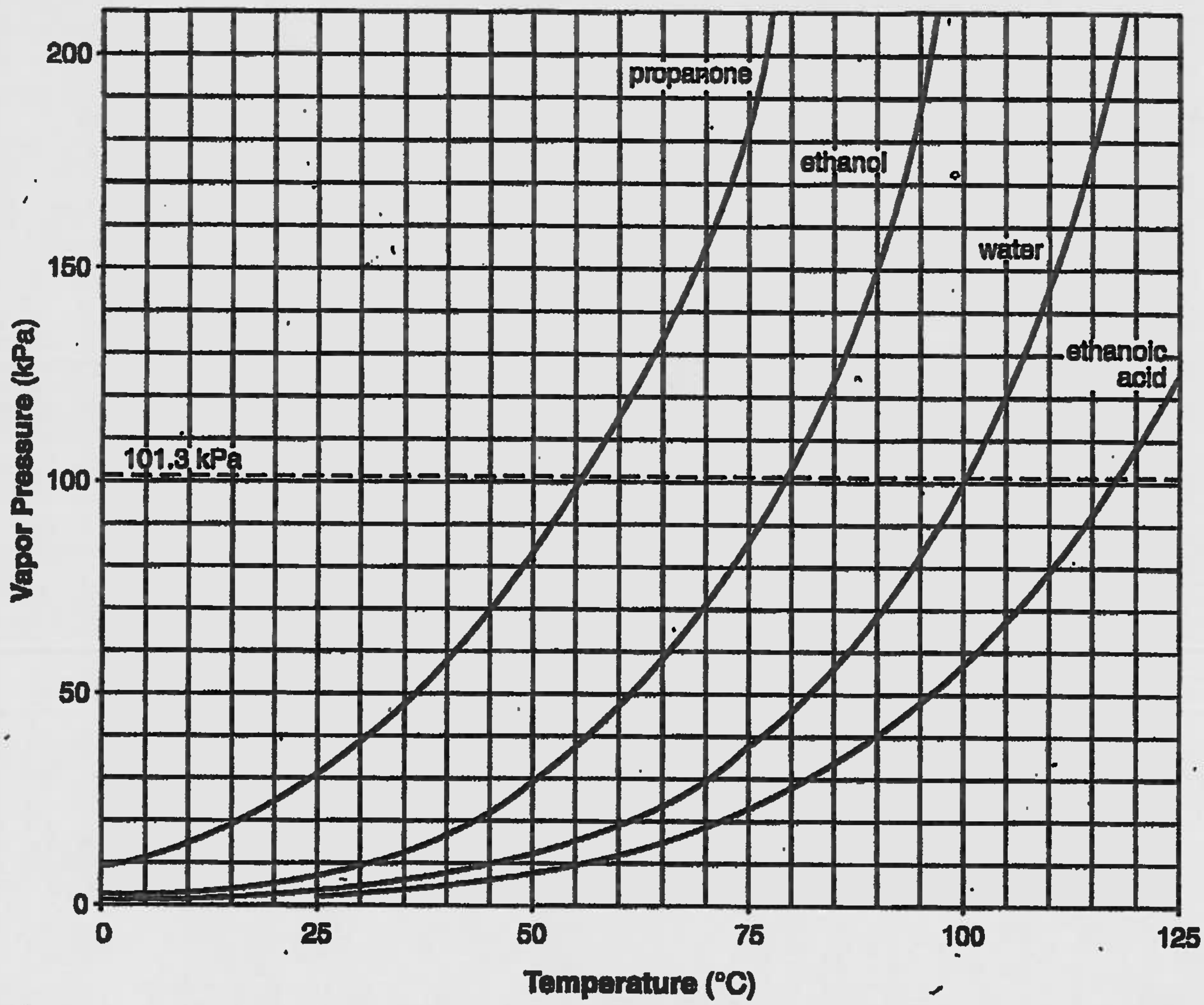
(1) 82 kPa
(2) 90 kPa
(3) 101.3 kPa
(4) 120 kPa

- The vapor pressure of water is 50 kPa. The temperature of this sample of water is

(1) 37°C
(2) 62°C
(3) 82°C
(4) 92°C

- Four flasks are prepared, each containing 100mL of aqueous solutions of equal concentration. Flask A contains $KCl(aq)$, flask B contains $CH_3OH(aq)$, flask C contains $Ba(OH)_2(aq)$ and flask D contains $CH_3COOH(aq)$. Which solution has the lowest freezing point? Explain your answer.

Table H — Vapor Pressure of Four Liquids



verview:

liquid is the form of matter that has definite volume but no definite shape. A liquid takes the shape of e container it is in. Above the surface of a liquid, there is always found the gaseous form of that liquid, alled a vapor. The term vapor refers to the gas phase of a substance that is ordinarily a solid or liquid t that temperature. This vapor above the surface of a liquid exerts a characteristic pressure called vapor ressure.

he Table:

his table shows the vapor pressure, in kPa, of four liquids as a function of temperature. The graph shows hat propanone has the greatest vapor pressure at any given temperature compared to the other three iq , while ethanoic acid has the lowest vapor pressure at any given temperature compared to the other hre liquids.

To determine the vapor pressure of a liquid at a specific temperature, move directly up from the given temperature until you reach the intersection point of the liquid's vapor pressure curve. Reading across to the vapor pressure axis gives the vapor pressure of that liquid at that temperature. The dotted horizontal line labeled 101.3 kPa is standard pressure (see Table A).

Temperature vs. Vapor Pressure
As the temperature increases, the vapor pressure increases. This is due to an increased amount of vapor and the greater average kinetic energy of the vapor particles. As the pressure on the surface of a liquid increases, the boiling point of the liquid increases. This is caused by the need to reach a higher vapor pressure to equal the increased pressure on the surface of the liquid.

Boiling Point and Vapor Pressure
The boiling point of a liquid is the temperature at which the vapor pressure is equal to the atmospheric pressure on the surface of the liquid. Therefore, when a liquid is boiling, the atmospheric pressure on the liquid can be read from the vapor pressure axis since they are equal to each other. When the atmospheric pressure is equal to standard pressure, the boiling point is called the normal boiling point. Reading from the graph at standard pressure (101.3 kPa), the normal boiling points of propanone, ethanol, water and ethanoic acid are 56°C, 79°C, 100°C and 117°C, respectively.

Intermolecular Attraction
A higher boiling point for a liquid indicates a greater attraction between the molecules of that liquid. The vapor pressure curves on Table H indicate that propanone has the weakest intermolecular attraction and ethanoic acid has the greatest intermolecular attraction.

Additional Information:

- The vapor pressure depends only upon the nature of the liquid and the temperature. It does not depend upon the amount of liquid.

- If a temperature-pressure point lies on one of the vapor pressure curves, the liquid is boiling, changing from the liquid to the gas phase. If the intersection point of the temperature and atmospheric pressure (read from the vapor pressure axis) of the substance is to the left of its vapor pressure curve, that substance is a liquid. If the intersection point lies to the right of the vapor pressure curve, it is a gas. For example, at 25°C and 150 kPa pressure, propanone is in the liquid phase.

Set 1 — Vapor Pressure of Four Liquids

.ich substance has the lowest vapor pressure at 75°C?

(1) water
(2) ethanoic acid
(3) propanone
(4) ethanol

1 _____

According to Reference Table H, what is the vapor pressure of propanone at 45°C?

(1) 22 kPa (3) 70. kPa
(2) 33 kPa (4) 98 kPa

2 _____

: boiling point of a liquid is the temperature at which the vapor pressure of the liquid is equal to the pressure on the surface of the liquid. What is the boiling point of propanone if the pressure on its surface is 48 kilopascals?

(1) 25°C (3) 35°C
(2) 30.°C (4) 40.°C

3 _____

4. As the temperature of a liquid increases, its vapor pressure

(1) decreases
(2) increases
(3) remains the same

4 _____

5. As the pressure on the surface of a liquid decreases, the temperature at which the liquid will boil

(1) decreases
(2) increases
(3) remains the same

5 _____

6. Using your knowledge of chemistry and the information in Reference Table H, which statement concerning propanone and water at 50°C is true?

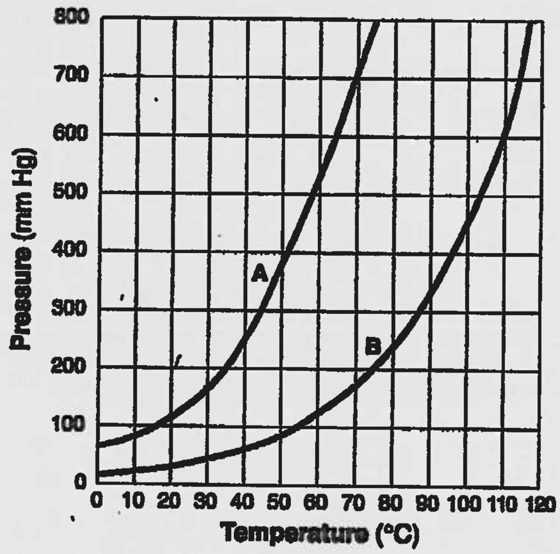
(1) Propanone has a higher vapor pressure and stronger intermolecular forces than water.
(2) Propanone has a higher vapor pressure and weaker intermolecular forces than water.
(3) Propanone has a lower vapor pressure and stronger intermolecular forces than water.
(4) Propanone has a lower vapor pressure and weaker intermolecular forces than water.

6 _____

. A liquid boils when the vapor pressure of the liquid equals the atmospheric pressure on the surface of the liquid. Using Reference Table H, determine the boiling point of water when the atmospheric pressure is 90. kPa.

____

Base your answers to question 8 using your knowledge of chemistry and on the graph below, which shows the vapor pressure curves for liquids *A* and *B*. **Note:** The pressure is given in mm Hg – millimeters of mercury.



8. *a*) What is the vapor pressure of liquid *A* at 70°C? Your answer must include correct units.

____

b) At what temperature does liquid *B* have the same vapor pressure as liquid *A* at 70°C? Your answer must include correct units.

____

c) At 400 mm Hg, which liquid would reach its boiling point first? ____________________

d) Which liquid will evaporate more rapidly? Explain your answer in terms of intermolecular forces.

____

____

____

vapor pressure

1. Which liquid has the lowest vapor pressure at 65°C?
1) ethanoic acid 2) ethanol 3) propanone 4) water

2. Which liquid has the highest vapor pressure at 75°C?
1) ethanoic acid 2) ethanol 3) propanone 4) water

3. Using your knowledge of chemistry and the information in Reference Table *H*, which statement concerning propanone and water at 50°C is true?
1) Propanone has a higher vapor pressure and stronger intermolecular forces than water.
2) Propanone has a higher vapor pressure and weaker intermolecular forces than water.
3) Propanone has a lower vapor pressure and stronger intermolecular forces than water.
4) Propanone has a lower vapor pressure and weaker intermolecular forces than water.

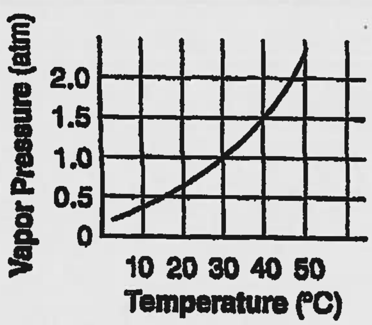
4. According to Reference Table *H*, what is the vapor pressure of propanone at 45°C?
1) 22 kPa 2) 33 kPa 3) 70 kPa 4) 98 kPa

5. As the temperature of a liquid increases, its vapor pressure
1) decreases 2) increases 3) remains the same

6. Which sample of water has the *lowest* vapor pressure?
1) 100 mL at 50°C 2) 200 mL at 30°C 3) 300 mL at 40°C 4) 400 mL at 20°C

7. Based on Reference Table *H*, which substance has the weakest intermolecular forces?
1) ethanoic acid 2) ethanol 3) propanone 4) water

8. The graph below shows the relationship between vapor pressure and temperature for substance *X*.



What is the normal boiling point for substance *X*?
1) 50°C 2) 20°C 3) 30°C 4) 40°C

9. When the vapor pressure of water is 30 kPa, the temperature of the water is
1) 20°C 2) 40°C 3) 70°C 4) 100°C

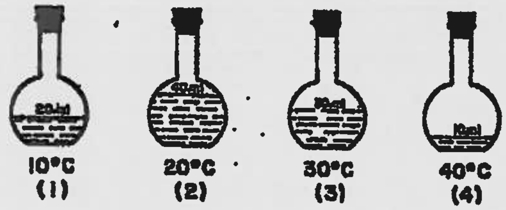
10. In a closed system, as the temperature of a liquid increases, the vapor pressure of the liquid
1) decreases 2) increases 3) remains the same

11. When the vapor pressure of a liquid is equal to the atmospheric pressure, the liquid will
1) freeze 2) boil 3) melt 4) condense

12. If the pressure on the surface of water in the liquid state is 47 kPa, the water will boil at
1) 0.0°C 2) 40°C 3) 80°C 4) 101.3°C

13. What is the vapor pressure of a liquid at its normal boiling temperature?
1) 1 kPa 2) 101.3 kPa 3) 273 kPa 4) 760 kPa

14. Base your answer to the following question on the diagrams below of four sealed flasks, each of which contains $H_2O(\ell)$ at the temperature shown.

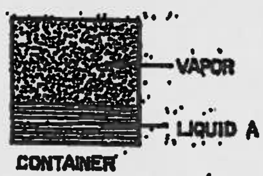


In which flask is the equilibrium vapor pressure of water the greatest?
1) 1 2) 2 3) 3 4) 4

15. As the atmospheric pressure decreases, the temperature at which water will boil in an open container
1) decreases 2) increases 3) remains the same

16. The vapor pressure of 25 milliliters of water at 25°C will be the same as
1) 50. ml of water at 25°C 2) 25 ml of water at 50°C 3) 50. ml of alcohol at 25°C 4) 25 ml of alcohol at 50°C

17.



Liquid *A* is confined in a container as shown in the diagram above. The equilibrium vapor pressure of liquid *A* depends on the
1) amount of vapor in the container
2) amount of liquid in the container
3) temperature of liquid *A*
4) size of the confining container

1. Compared to pure water, an aqueous solution of calcium chloride has a
   (1) higher boiling point and higher freezing point
   (2) higher boiling point and lower freezing point
   (3) lower boiling point and higher freezing point
   (4) lower boiling point and lower freezing point

2. As a solute is added to a solvent, what happens to the freezing point and the boiling point of the solution?
   (1) The freezing point decreases and the boiling point decreases.
   (2) The freezing point decreases and the boiling point increases.
   (3) The freezing point increases and the boiling point decreases.
   (4) The freezing point increases and the boiling point increases.

3. When ethylene glycol (an antifreeze) is added to water, the boiling point of the water
   (1) decreases, and the freezing point decreases
   (2) decreases, and the freezing point increases
   (3) increases, and the freezing point decreases
   (4) increases, and the freezing point increases

4. What occurs as a salt dissolves in water?
   (1) The number of ions in the solution decreases, and the freezing point decreases.
   (2) The number of ions in the solution decreases, and the freezing point increases.
   (3) The number of ions in the solution increases, and the freezing point decreases.
   (4) The number of ions in the solution increases, and the freezing point increases.

5. Which solution containing 1 mole of solute dissolved in 1000 grams of water has the lowest freezing point?
   (1) $KOH(aq)$ (3) $C_2H_5OH(aq)$
   (2) $C_2H_{12}O_6(aq)$ (4) $C_{12}H_{22}O_{11}(aq)$

6. Of the following solutions, the one that will boil at the highest temperature contains 1 mole of nonvolatile solute dissolved in
   (1) 250 g of solvent (3) 750 g of solvent
   (2) 500 g of solvent (4) 1,000 g of solvent

7. Which solution has the highest boiling point?
   (1) 1.0 M KNO_3 (3) 1.0 M $Ca(NO_3)_2$
   (2) 2.0 M KNO_3 (4) 2.0 M $Ca(NO_3)_2$

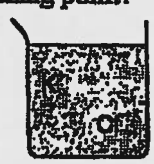
8. Which solute, when added to 1,000 grams of water, will produce a solution with the highest boiling point?
   (1) 29 g of NaCl (3) 31 g of $C_2H_5O_2$
   (2) 58 g of NaCl (4) 62 g of $C_2H_6O_2$

9. Which concentration of a solution of CH_3OH in water has the *lowest* freezing point?
   (1) 0.1 M (3) 0.001 M
   (2) 0.01 M (4) 0.0001 M

10. Which of the following solutions, each containing a nonvolatile solute, will boil at the highest temperature?
   (1) 1 mole of electrolyte dissolved in 1000 g of H_2O
   (2) 2 moles of electrolyte dissolved in 1000 g of H_2O
   (3) 1 mole of nonelectrolyte dissolved in 1000 g of H_2O
   (4) 2 moles of nonelectrolyte dissolved in 1000 g of H_2O

11. Which aqueous solution has the *lowest* freezing point?
   (1) 1.0 M $C_6H_{12}O_6$ (3) 1.0 M CH_3COOH
   (2) 1.0 M C_2H_5OH (4) 1.0 M NaCl

12. Which 1-molal aqueous solution has the *lowest* freezing point?

(1)


(3)


(2)


(4)


13. A 1 molal solution of $MgCl_2$ has a higher boiling point than a 1 molal solution of
   (1) $FeCl_3$ (3) $BaCl_2$
   (2) $CaCl_2$ (4) NaCl

salting Roads

THE SOLUTION FOR WINTER DRIVING

By Doris R. Kimbrough

Most of you are probably planning to be in a car at some point this winter, so be sure to wear your seatbelt, obey the speed limit, and be extra cautious in "winter driving conditions". It is actually physics that makes being in a car on an icy road hazardous; friction is an important part of keeping a car under control. However, your state and local transportation departments make use of some pretty interesting chemistry to keep roads safer for travelers in the wintertime. In addition to plowing, one of the ways that highway workers keep roads clear of ice and snow is by spreading salt on the roads. Even though salt can cause rust and corrosion on cars, bridges, and other parts of the highway, it more than makes up for this costly damage by saving lives. Let's take a look at how it works.

Freezing point depression

Pure water freezes at 0°C; adding salt to water depresses or lowers the freezing point below zero. When you remove heat from water (or any substance), the molecules slow down. The freezing process occurs when the molecules stop sliding and tumbling all over each other (liquid phase) and settle into fixed positions in a large network called a *crystal lattice*, which is the solid phase. The molecules are still moving, but in the solid phase that motion involves bonds stretching and compressing or the atoms wiggling a little bit. This is called vibrational motion.

When a solution of salt in water is cooled to a low enough temperature, the water molecules begin to stick together in an organized way to form solid crystals. The crystal framework tends not to include the salt ions because the ions would disturb this organization. So when you cool a solution enough, the ice crystals that start to form are made of pure frozen water. You can actually purify salty water like this, by freezing a portion of the solution and washing the salty water off of the ice crystals and then thawing the ice to produce pure water. Eventually, if you cooled it enough, the whole solution will freeze, but it does not have a sharply defined freezing temperature. Getting the water molecules organized into a crystal from a solution requires that you remove more energy (actually free energy) than if you are freezing pure water, so the water in a solution typically does not start freezing until it reaches a lower temperature than the normal freezing point. This is true of all solutions, not just those made with water.

The difference in temperature between where the pure solvent (water in our case) freezes and where the solution starts to freeze is called the *freezing point depression*. How low

you go depends upon what the solvent is and how concentrated the solution is. The more concentrated the solution, the lower the freezing point. This is why if you freeze a solution completely, you have to keep lowering the temperature. As the pure water freezes away from the solution, the concentration of the solution remaining increases. How does your Department of Transportation use this chemistry to keep roadways clear of snow and ice in the wintertime?

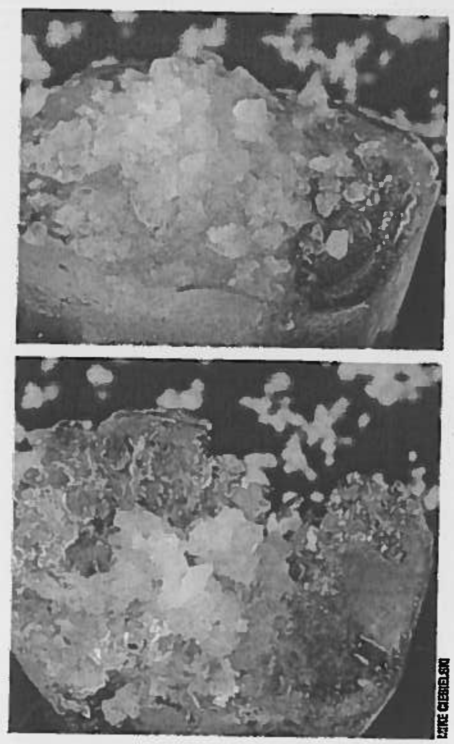
Sodium chloride and calcium chloride.

Salting roads

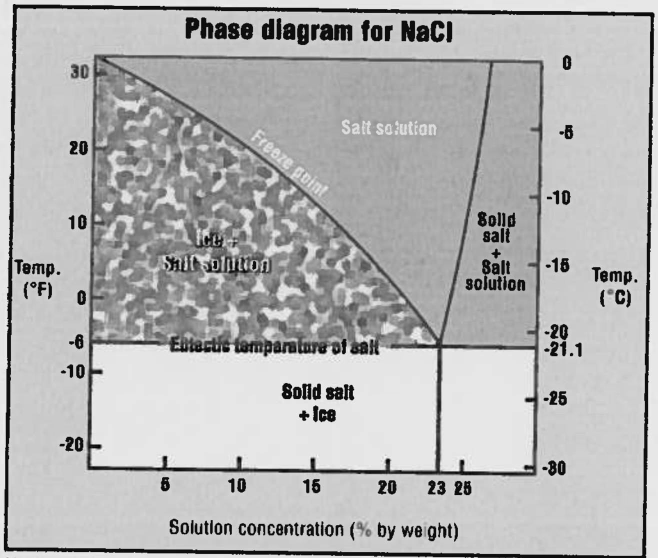
Highway workers use salt in two ways: 1) to melt ice that is already on the roadway and 2) to prevent ice from forming on the roadway. The second one is a little easier to understand, so we will start there. Let's say a snowstorm is forecast for your town. Municipal workers get out and spread salt on the roads. As the storm hits, snow starts to fall, but the road surface is warmer than the air, so the first flakes melt. As they melt, the salt dissolves in the liquid water. Now you have a solution of salty water, which has a lower freezing point than pure snow, so that even though the additional snow might cool the road enough to "stick" to the road surface, the temperature will not get cold enough to freeze the solution the way it could freeze pure water. In the end, the real snow removal is done by the plows, but salt plays a crucial role in preventing snow and ice from bonding to the pavement.

"Hold it!" you say. Suppose the temperature does get cold enough to freeze the solution. Or suppose that enough snow falls so that the salt water solution gets too dilute to work, what happens then? In both of those cases, snow could build up on the road. This is why your highway workers are out there round-the-clock, plowing and spreading more salt as long as the storm is in the area. Communities with really cold temperatures, like parts of Canada and Alaska, where it can get to –20°C (below zero on the Fahrenheit scale), don't even bother with salting the roads, because it doesn't help. They typically plow off as much as they can and use gravel or sand to add traction.

How does it work if the road already has snow or ice on it? You may have seen how spreading salt on an icy sidewalk will cause the snow or ice to melt. The very beginning of the melting process begins where solid ice contacts solid salt. At the surface of ice crystals, water molecules do not have the same stable arrangement they have on the interior; they are more mobile and more reactive. So the surface ice reacts with the surface of the salt crystals, allowing a small amount of salty



Water freezes at lower temperatures when it is saltier. But solubility of salt in water decreases with decreasing temperature. So at some point, a solution is too cold to hold the salt that keeps it from freezing. The *eutectic* temperature is the lowest temperature at which a mixture of two or more substances can stay liquefied. For a NaCl and water mixture, the eutectic temperature is –21.1 °C. For road ice, –10 °C is the practical limit for salt.



This phase diagram illustrates the impact of NaCl concentration and temperature on the phase of an aqueous NaCl solution.

solution to form. This first step is relatively slow, but then the growing solution continues to dissolve more salt and melt more ice. Passing vehicles warm the slush through friction, which speeds the dissolution of the ice and may crush the salt and ice together, which will increase the surface areas of the particles in contact with each other. Some communities use "prewetted" salt (usually rock salt with a $CaCl_2$ solution sprayed on it) to speed the process.

Road salt history, fun facts, and current technology

Putting salt on the roads to lessen the buildup of snow and ice began in the 1930s, and by the 1960s, it was used by most communities where snow and ice are a problem. Concerns about the effect of the use of sodium chloride (common table salt) on the environment have prompted some state and local road crews to explore the use of more environmentally friendly salts, such as magnesium chloride and calcium chloride. These two salts have the advantage of being more effective at lowering the freezing point and there is some interesting chemistry behind this benefit.

As you probably know, sodium chloride has the formula NaCl. When it dissolves in water, it dissociates into its ions: Na^+ and Cl^-, producing two ions in solution for every NaCl formula unit. Magnesium chloride ($MgCl_2$) and calcium chloride ($CaCl_2$) dissociate to three ions each because the metal has a 2^+ charge and there are two chlorides per metal ion:

$$MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$$
$$CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$$

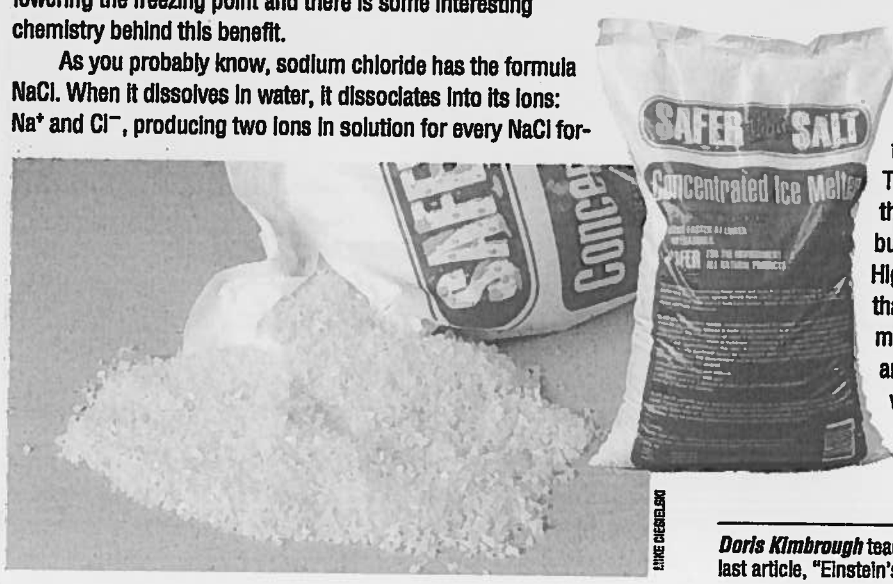
It is the number of dissolved particles that determine the extent of the lowering of the freezing point of a solution. So although NaCl produces two ions, $MgCl_2$ and $CaCl_2$ each produce three, making them more effective. Other variations include mixtures of the magnesium and calcium chlorides, as well as magnesium and calcium acetates, $Mg(C_2H_3O_2)_2$ and $Ca(C_2H_3O_2)_2$.

	Practical Melt Temp.	Eutectic Temp.
$CaCl_2$	−32°C	−56°C
$MgCl_2$	−15°C	−33°C

The technology of salting roads has become fairly sophisticated. Often, these salts are dissolved in water or some other solvent so that they can be sprayed onto the road surface. Having the deicing substance in a solution (i.e., fluid) form makes it possible to pump through hoses, which allows for a more targeted application. In addition, various anticorrosive substances are added to protect highways and cars from the damage the salts can cause over time. Some 15 million tons of deicing salt is used each year in the United States and about 4–5 million tons in Canada.

You may have seen signs that warn about bridges freezing before road surfaces. This is because bridges are more exposed and not insulated by the ground from underneath like the rest of the highway. Some high-tech highway bridges have been constructed with deicing sprayers built right into the pavement, complete with sensors that detect when conditions are right (e.g., cold temperatures, high wind speeds, high humidity) for ice to form. The sensors detect the possible formation of ice, and the deicing sprayers go to work to keep the roadway from freezing.

Highway engineers have been working with other interesting variations. One deicing material that is currently on the market mixes magnesium chloride with sugar cane or sugar beet molasses. The sticky molasses keeps the magnesium chloride from getting blown or washed off the road surface. There are also substances that are added directly to the top layer of concrete or asphalt when the road is built or repaved that help prevent ice from forming. Highway workers can then get away with using less salt than before, which is cheaper, easier on the environment, and helpful in preventing corrosion. Scientists and engineers continue to develop new ways to keep winter highways safe while minimizing expense and environmental harm. Just another way that chemistry is keeping you out of harm's way.



LUKE CIESIELSKI

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Name____ Date________________

I n

Regents Chemistry

Salting Roads: The Solution for Winter Driving

1. What are pro's and con's to placing salt on the roads when it snows?

2. What does adding salt to water do to the freezing point of the water?

3. What happens when a substance freezes?

4. How can water be purified by freezing it?

5. What affects the freezing point of a solution?

6. What are the two ways highway workers use salt?

7. How does salt prevent ice from forming on the roadway?

8. What do cites with really cold temperatures do when it snows?

9. How does salt affect roads that already have snow on them?

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10. Why are magnesium chloride and calcium chloride more effective?

11. Why do bridges freeze before road surfaces?

12. What are some other ways highway engineers hinder roads from freezing in the cold weather?

Name______________________________ Date____________________

Solutions Review

1. According to Reference Table G, how does a decrease in temperature from 40°C to 20°C affect the solubility of NH $_3$ and KCl?
   1) The solubility of NH_3 decreases, and the solubility of KCl decreases.
   2) The solubility of NH_3 decreases, and the solubility of KCl increases.
   3) The solubility of NH_3 increases, and the solubility of KCl decreases.
   4) The solubility of NH_3 increases, and the solubility of KCl increases.

2. A solution contains 100 grams of a nitrate salt dissolved in 100 grams of water at 50°C. The solution could be a
   1) supersaturated solution of $NaNO_3$
   2) saturated solution of $NaNO_3$
   3) supersaturated solution of KNO_3
   4) saturated solution of KNO_3

3. What is the total number of moles of solute contained in 0.50 liter of 3.0 M HCl?
   1) 1.0 3) 3.0
   2) 1.5 4) 3.5

4. What is the concentration expressed in parts per million of a solution containing 15.0 grams of KNO_3 in 65.0 grams of H_2O?
   1) 1.88×10^5 ppm 3) 2.31×10^5 ppm
   2) 2.00×10^5 ppm 4) 5.33×10^6 ppm

5. Which concentration of a solution of CH_3OH in water has the *lowest* freezing point?
   1) 0.1 M 3) 0.001 M
   2) 0.01 M 4) 0.0001 M

6. What occurs as a salt dissolves in pure water?
   1) The number of ions in the solution decreases, and the freezing point decreases.
   2) The number of ions in the solution decreases, and the freezing point increases.
   3) The number of ions in the solution increases, and the freezing point decreases.
   4) The number of ions in the solution increases, and the freezing point increases.

7. When ethylene glycol (an antifreeze) is added to water, the boiling point of the water
   1) decreases, and the freezing point decreases
   2) decreases, and the freezing point increases
   3) increases, and the freezing point decreases
   4) increases, and the freezing point increases

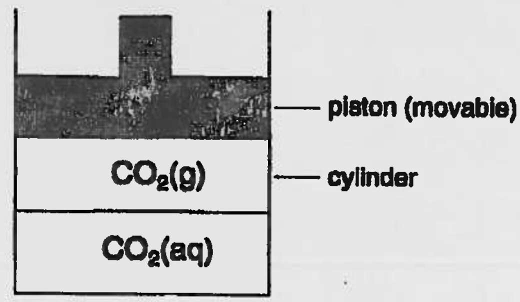
8. Which compound is most soluble in water?
   1) silver acetate 3) silver nitrate
   2) silver chloride 4) silver sulfate

9. Based on Reference Table F, which salt is *least* soluble?
   1) $FeCO_3$ 3) $BaCl_2$
   2) Na_2CO_3 4) $CaCl_2$

10. As the temperature increases from 0°C to 25°C the amount of NH_3 that can be dissolved in 100 grams of water
    1) decreases by 10 grams 3) increases by 10 grams
    2) decreases by 40 grams 4) increases by 40 grams

11. According to Reference Table G, how many grams of KNO $_3$ would be needed to saturate 200 grams of water at 70°C?
    1) 43 g 3) 134 g
    2) 86 g 4) 268 g

12. Given the diagram below that shows carbon dioxide in an equilibrium system at a temperature of 298 K and a pressure of 1 atm:



Which changes *must* increase the solubility of the carbon dioxide?
    1) increase pressure and decrease temperature
    2) increase pressure and increase temperature
    3) decrease pressure and decrease temperature
    4) decrease pressure and increase temperature

13. A solution contains 70 grams of $NaNO_3$ in 100 grams of water at 10°C. How many additional grams of $NaNO_3$ are required to saturate this solution?
    1) 10 3) 60
    2) 20 4) 70

14. Based on Reference Table G, when 100 grams of water saturated with KNO_3 at 70°C is cooled to 25°C, the total number of grams of KNO_3 that will precipitate is
    1) 40 3) 80
    2) 45 4) 95

15. A solution containing 90. grams of KNO_3 per 100. grams of H_2O at 50.°C is considered to be
    1) dilute and unsaturated
    2) dilute and supersaturated
    3) concentrated and unsaturated
    4) concentrated and supersaturated

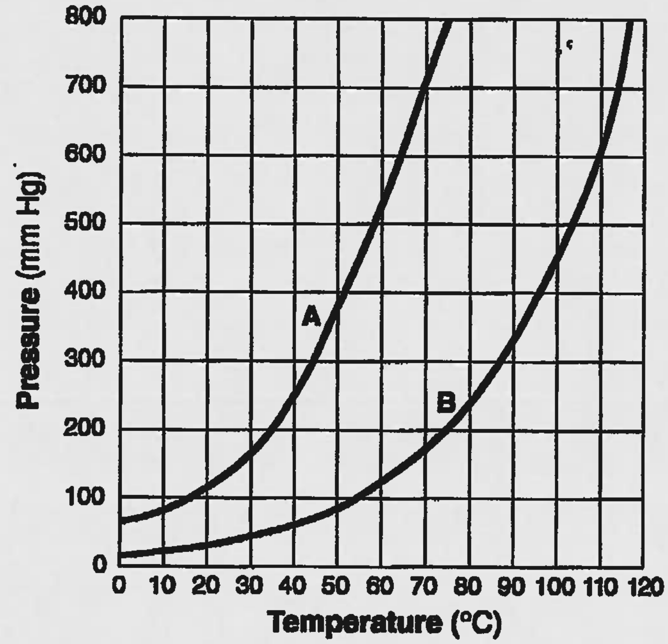
16. Which solution is most concentrated?
1) 0.1 mole of solute dissolved in 400 ml of solvent
2) 0.2 mole of solute dissolved in 300 ml of solvent
3) 0.3 mole of solute dissolved in 200 ml of solvent
4) 0.4 mole of solute dissolved in 100 ml of solvent

17. What is the concentration of a solution, in parts per million, if 0.02 gram of Na_3PO_4 is dissolved in 1000 grams of water?
1) 20 ppm
2) 2 ppm
3) 0.2 ppm
4) 0.02 ppm

18. As a solute is added to a solvent, what happens to the freezing point and the boiling point of the solution?
1) The freezing point decreases and the boiling point decreases.
2) The freezing point decreases and the boiling point increases.
3) The freezing point increases and the boiling point decreases.
4) The freezing point increases and the boiling point increases.

19. Show a correct numerical setup for determining how many liters of a 1.2 M solution can be prepared with 0.50 mole of $C_6H_{12}O_6$.

Base your answers to questions 20 and 21 on the graph below, which shows the vapor pressure curves for liquids *A* and *B*.



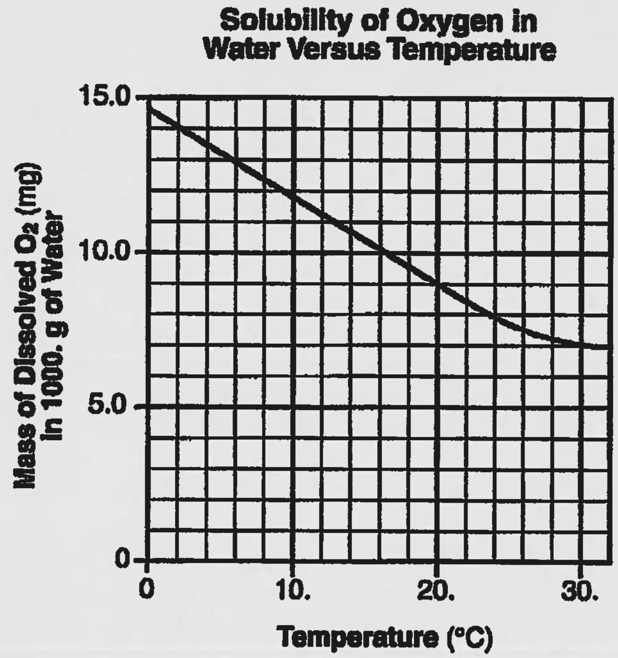
20. What is the vapor pressure of liquid *A* at 70°C? Your answer must include correct units.

21. Which liquid will evaporate more rapidly? Explain your answer in terms of intermolecular forces.

____

Base your answers to questions 22 and 23 on the information below

Scientists who study aquatic ecosystems are often interested in the concentration of dissolved oxygen in water. Oxygen, O2, has a very low solubility in water, and therefore its solubility is usually expressed in units of milligrams per 1000. grams of water at 1.0 atmosphere. The graph below shows a solubility curve of oxygen in water.



22. A student determines that 8.2 milligrams of oxygen is dissolved in a 1000.-gram sample of water at 15°C and 1.0 atmosphere. In terms of saturation, what type of solution is this sample?

23. Explain, in terms of molecular polarity, why oxygen gas has low solubility in water. Your response must include *both* oxygen and water.

Base your answers to questions 24 and 25 on the properties of propanone.

24. Explain, in terms of molecular energy, why the vapor pressure of propanone increases when its temperature increases.

25. A liquid's boiling point is the temperature at which its vapor pressure is equal to the atmospheric pressure. Using Reference Table *H*, what is the boiling point of propanone at an atmospheric pressure of 70 kPa?

26. A liquid boils when the vapor pressure of the liquid equals the atmospheric pressure on the surface of the liquid. Using Reference Table *H*, determine the boiling point of water when the atmospheric pressure is 90. kPa.

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27. Base your answer to the following question on the information and table below.

A student conducts an experiment to determine how the temperature of water affects the rate at which an antacid tablet dissolves in the water. The student has three antacid tablets of the same size and composition. The student drops one tablet into each of three beakers containing 200. milliliters of water at different temperatures and measures the time it takes for each tablet to completely dissolve. The results are shown in the table below.

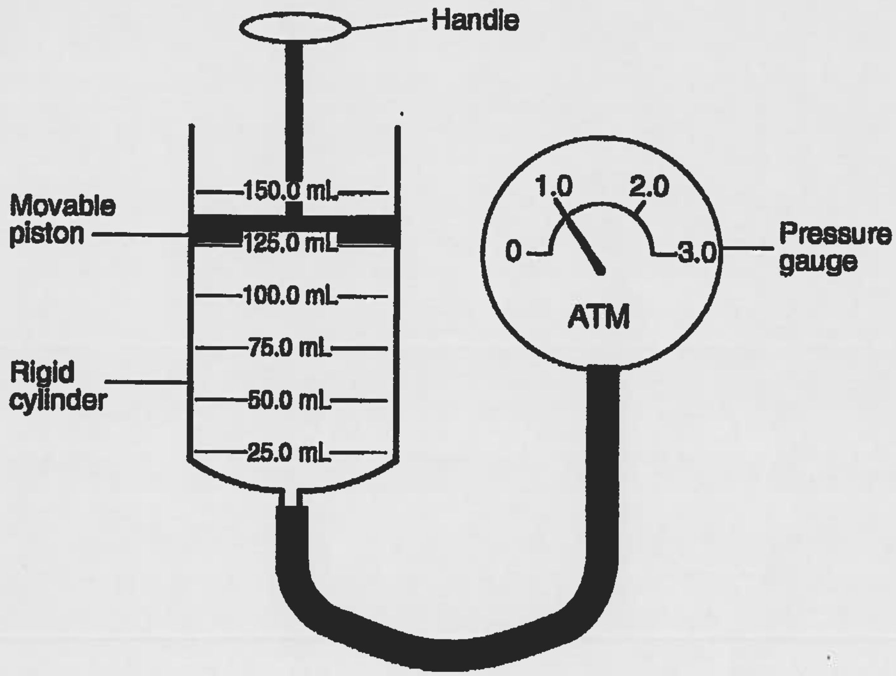
Dissolving Data for Three Antacid Tablets

Beaker	Original Temperature of Water (°C)	Time for Tablet to Dissolve (s)
1	20.	40.
2	30.	25
3	40.	10.

Describe the effect of water temperature on the rate of dissolving.

28. **Base your answer to the following question on the information below.**

A rigid cylinder is fitted with a movable piston. The cylinder contains a sample of helium gas, He(g), which has an initial volume of 125.0 milliliters and an initial pressure of 1.0 atmosphere, as shown below. The temperature of the helium gas sample is 20.0°C.



Express the initial volume of the helium gas sample, in liters.

Name ______________________________ Date ______________

Solutions Review

____1. The solubility of $KClO_3(s)$ in water increases as the

1) temperature of the solution increases
2) temperature of the solution decreases
3) pressure on the solution increases
4) pressure on the solution decreases

____2. Solubility data for four different salts in water at 60°C are shown in the table below.

Salt	Solubility in Water at 60°C
A	10 grams / 50 grams H_2O
B	20 grams / 60 grams H_2O
C	30 grams / 120 grams H_2O
D	40 grams/ 80 grams H_2O

Which salt is most soluble at 60°C?

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

____3. According to Reference Table *G*, which of these substances is most soluble at 60°C?

1) NaCl 3) $KClO_3$
2) KCl 4) NH_4Cl

____4. According to Reference Table *G*, how many grams of KNO_3 would be needed to saturate 200 grams of water at 70°C?

1) 43 g 3) 134 g
2) 86 g 4) 268 g

____5. As the pressure on a gas confined above a liquid increases, the solubility of the gas in the liquid

1) decreases 3) remains the same
2) increases

____6. What is the maximum number of grams of NH_4Cl that will dissolve in 200 grams of water at 70°C?

1) 60 3) 100
2) 70 4) 120

____7. The solubility of a salt in a given volume of water depends primarily on the

1) surface area of the salt crystals
2) temperature of the water
3) rate at which the salt and water are stirred
4) pressure on the surface of the water

____8. According to Reference Table G, how does a decrease in temperature from 40°C to 20°C affect the solubility of NH_3 and KCl?

1) The solubility of NH_3 decreases, and the solubility of KCl decreases.
2) The solubility of NH_3 decreases, and the solubility of KCl increases.
3) The solubility of NH_3 increases, and the solubility of KCl decreases.
4) The solubility of NH_3 increases, and the solubility of KCl increases.

____9. A student determined the mass, in grams, of compound *X* that would saturate 30. grams of water over a temperature range of 40.°C in 10.-degree intervals. The results are tabulated below.

Grams of Dissolved Compound *X*	Temperature of 30. grams of H_2O
2.0 g	10.°C
4.0 g	20.°C
8.0 g	30.°C
16 g	40.°C
32 g	50.°C

If this solubility trend continues, what is the total number of grams of compound *X* that will dissolve in 30. grams of water at 60.°C?

1) 16 3) 48
2) 32 4) 64

____10. As additional $KNO_3(s)$ is added to a saturated solution of KNO_3 at constant temperature, the concentration of the solution

1) decreases 3) remains the same
2) increases

____11. Which phrase describes the molarity of a solution?

1) liters of solute per mole of solution
2) liters of solution per mole of solution
3) moles of solute per liter of solution
4) moles of solution per liter of solution

____12. How many total moles of KNO_3 must be dissolved in water to make 1.5 liters of a 2.0 M solution?

1) 0.50 mol 3) 3.0 mol
2) 2.0 mol 4) 1.3 mol

____13. What is the freezing point of a solution that contains 1.00 mole of a nonelectrolyte dissolved in 1,000. grams of water?

1) 0.00°C 3) −1.86°C
2) 0.520°C 4) −3.72°C

____14. Which solution is most concentrated?

1) 0.1 mole of solute dissolved in 400 ml of solvent
2) 0.2 mole of solute dissolved in 300 ml of solvent
3) 0.3 mole of solute dissolved in 200 ml of solvent
4) 0.4 mole of solute dissolved in 100 ml of solvent

____15. A student wants to prepare a 1.0-liter solution of a specific molarity. The student determines that the mass of the solute needs to be 30. grams. What is the proper procedure to follow?

1) Add 30. g of solute to 1.0 L of solvent.
2) Add 30. g of solute to 970. mL of solvent to make 1.0 L of solution.
3) Add 1000. g of solvent to 30. g of solute.
4) Add enough solvent to 30. g of solute to make 1.0 L of solution.

____16. What is the concentration of $O_2(g)$, in parts per million, in a solution that contains 0.008 gram of $O_2(g)$ dissolved in 1000. grams of $H_2O(l)$?

1) 0.8 ppm 3) 80 ppm
2) 8 ppm 4) 800 ppm

____17. Compared to a 0.1 M aqueous solution of NaCl, a 0.8 M aqueous solution of NaCl has a

1) higher boiling point and a higher freezing point
2) higher boiling point and a lower freezing point
3) lower boiling point and a higher freezing point
4) lower boiling point and a lower freezing point

____18. At standard pressure when NaCl is added to water, the solution will have a

1) higher freezing point and a lower boiling point than water
2) higher freezing point and a higher boiling point than water
3) lower freezing point and a higher boiling point than water
4) lower freezing point and a lower boiling point than water

____19. Which solution has the highest boiling point?

1) 1.0 M KNO_3 3) 1.0 M $Ca(NO_3)_2$
2) 2.0 M KNO_3 4) 2.0 M $Ca(NO_3)_2$

____20. Which compound is insoluble in water?

1) $BaSO_4$ 3) $KClO_3$
2) $CaCrO_4$ 4) Na_2S

____21. According to Reference Table F, which of these compounds is most soluble at 298 K and 1 atm?

1) $PbCl_2$ 3) $PbCrO_4$
2) AgCl 4) $PbCO_3$

____22. A student obtained the following data in determining the solubility of a substance.

Temperature (°C)	Grams of Solute/ 100 g H_2O
10	70
30	45
60	23
90	11

Which graph best represents the solubility curve drawn from the results obtained by the student?

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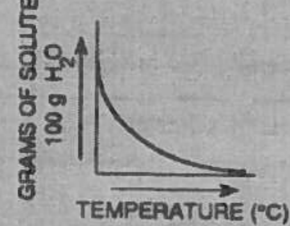

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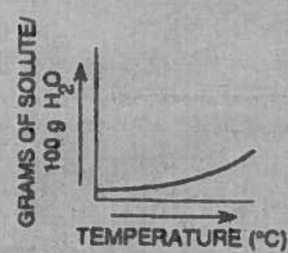

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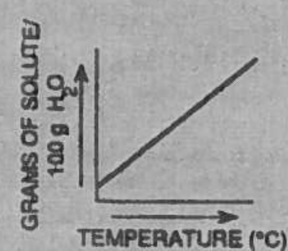

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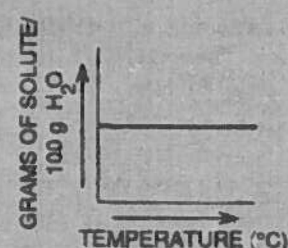

____23. A student prepares four aqueous solutions, each with a different solute. The mass of each dissolved solute is shown in the table below.

Mass of Dissolved Solute for Four Aqueous Solutions

Solution Number	Solute	Mass of Dissolved Solute (per 100. g of H_2O at 20.°C)
1	KI	120. g
2	$NaNO_3$	88 g
3	KCl	25 g
4	$KClO_3$	5 g

Which solution is saturated?

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

____24. According to your Reference Tables, which substance forms an unsaturated solution when 80 grams of the substance is dissolved in 100 grams of H_2O at 10°C?

1) KI
3) $NaNO_3$
2) KNO_3
4) NaCl

____25. Based on Reference Table *G*, what is the maximum number of grams of KCl(s) that will dissolve in 200 grams of water at 50°C to produce a saturated solution?

1) 38 g
3) 58 g
2) 42 g
4) 84 g

____26. Based on Reference Table G, what change will cause the solubility of $KNO_3(s)$ to increase?

1) decreasing the pressure
2) increasing the pressure
3) decreasing the temperature
4) increasing the temperature

____27. When 5 grams of KCl are dissolved in 50. grams of water at 25°C, the resulting mixture can be described as

1) heterogeneous and unsaturated
2) heterogeneous and supersaturated
3) homogeneous and unsaturated
4) homogeneous and supersaturated

____28. A solution containing 90. grams of KNO_3 per 100. grams of H_2O at 50.°C is considered to be

1) dilute and unsaturated
2) dilute and supersaturated
3) concentrated and unsaturated
4) concentrated and supersaturated

____29. Which is a saturated solution?

1) 40 g NH_4Cl in 100 g water at 50°C
2) 2 g SO_2 in 100 g water at 10°C
3) 52 g KCl in 100 g water at 80°C
4) 120 g KI in 100 g water at 20°C

____30. What is the total number of moles of NaCl(s) needed to make 3.0 liters of a 2.0 M NaCl solution?

1) 6.0 mol
3) 1.0 mol
2) 8.0 mol
4) 0.70 mol

____31. What is the molarity of a solution that contains 0.50 mole of NaOH in 0.50 liter of solution?

1) 1.0 M
3) 0.25 M
2) 2.0 M
4) 0.50 M

____32. Which preparation produces a 2.0 M solution of $C_6H_{12}O_6$? [molecular mass = 180.0]

1) 90.0 g of $C_6H_{12}O_6$ dissolved in 500.0 mL of solution
2) 90.0 g of $C_6H_{12}O_6$ dissolved in 1000. mL of solution
3) 180.0 g of $C_6H_{12}O_6$ dissolved in 500.0 mL of solution
4) 180.0 g of $C_6H_{12}O_6$ dissolved in 1000. mL of solution

____33. If 0.025 gram of $Pb(NO_3)_2$ is dissolved in 100. grams of H_2O, what is the concentration of the resulting solution, in parts per million?

1) 2.5×10^{-4} ppm
3) 250 ppm
2) 2.5 ppm
4) 4.0×10^3 ppm

____34. How many grams of NaCl are needed to be dissolved in water to make 1.0 gram of a 100.0 ppm solution?

1) 1.0×10^{-4} g
3) 1.0×10^{-2} g
2) 1.0×10^{-3} g
4) 1.0×10^{-1} g

____35. Compared to a 2.0 M aqueous solution of NaCl at 1 atmosphere, a 3.0 M aqueous solution of NaCl at 1 atmosphere has a

1) lower boiling point and a higher freezing point
2) lower boiling point and a lower freezing point
3) higher boiling point and a higher freezing point
4) higher boiling point and a lower freezing point

36. Base your answer to the following question on the information below

An unsaturated solution is made by completely dissolving 20.0 grams of $NaNO_3$ in 100.0 grams of water at 20.0°C.

Determine the minimum mass of NaNO3 that must be added to this unsaturated solution to make a saturated solution at 20.0°C.

37. Base your answer to the following question on on the information below.

A solution is made by completely dissolving 90. grams of KNO_3(s) in 100. grams of water in a beaker. The temperature of this solution is 65°C.

Describe the effect on the solubility of KNO_3(s) in this solution when the pressure on the solution increases.

Name ______________________________ Date__________ Period______

Kool Aid Koncentration

Objective: Prepare different concentrations of Kool Aid using Molarity!

Safety Precautions: Wear safety goggles at all times during the lab

Procedure:

1. Assume that Kool Aid is 100% sugar, $C_{12}H_{22}O_{11}$
2. Calculate the amount of Kool Aid that is needed to make 100mL each of 0.010M, 0.10M and 1.0M solutions. SHOW THE CALCULATIONS USED TO DETERMINE THE MOLARITY IN THE CALCULATIONS SECTION
3. Make the solutions of each concentration using the paper cups and plastic spoons
4. Observe and record the color and taste of each solution.

Calculations:

1. Calculate the amount of Kool Aid needed to make 100mL of 0.010M solution

2. Calculate the amount of Kool Aid needed to make 100mL of 0.10M solution

3. Calculate the amount of Kool Aid needed to make 100mL of a 1.0M solution

Data:

Table 1:

Molarity	Mass of Kool Aid	Volume of H_2O	Solution Color	Taste of Solution

Conclusions:

1. **Summarize your results relating the concentration of the Kool Aid, the color and the taste.**

2. **Estimate what Molarity of Kool Aid would taste good to you. Explain how you came up with your answer.**

3. **Describe how the concentration would be calculated differently if it were percent by mass**